

AP Chemistry Syllabus for 2015-2016

Bethlehem Central High School

Mr. Charles Reed – teacher

Textbook: Chemistry: The Central Science (10th ed.; ©2006) by Brown, LeMay, & Bursten

Lab Manual: AP Chemistry Laboratory Manual @ Bethlehem Central HS (from a variety of sources)

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GRADING PROCEDURES – Mr. Reed: AP Chemistry for 2015 – 2016

Grades in this course are based on a point system. At the end of each quarter, all your earned points will be summed and then divided by the total number of points for that quarter in order to determine your grade.

GRADES WILL NOT BE GIVEN OUT UNTIL REPORT CARDS ARE DISTRIBUTED!

POINT VALUES:

Homework (on time – when collected, not a minute after) – variable points

Homework (late) – points will be deducted at the discretion of the teacher

- (1) Please use the heading at the top of your homework paper that is shown on the right.
- (2) Homework should be submitted on loose leaf notebook paper with a “smooth edge” (tear or cut off any “ruffled edges”).
- (3) Show all pertinent work and do not cram your work on the page
- (4) Use the rules of significant figures and show appropriate units.
- (5) Put a box around your answers!
- (6) Skip a line in between each question.

Quizzes (about 43 min long) – 30 points

Tests (about 90 min long) – 120 points

Semester Exams (about 3 hours long) – equivalent to a quarter grade

Lab Reports (on time – by 3:00 PM) – 20 points

Lab Reports (late) – points will be deducted at the discretion of the teacher

AP Chemistry	Name:
HW #. #	Date:

Work missed due to an absence will be promptly made up. You have TWO days for each day you are LEGALLY absent! **You will not be notified that you owe work!** It is YOUR responsibility to find out the work you missed AND to KNOW when it is due. If you “forget”, and the make up period expires, any missed work will AUTOMATICALLY receive no credit. This includes homeworks, quizzes, tests, and lab activities.

WORK MISSED DUE TO CUTTING CLASS WILL RECEIVE ZERO POINTS!

Labs missed due to an absence can be made up on MOST ANY DAY. It is YOUR responsibility to schedule time and make up lab as soon as possible!

Makeup lab (2:15 to 3:00) is a class. Not attending will be considered as cutting!

In rare cases with extenuating circumstances, you may receive an extension for missed work BUT only if one is arranged in advance! **Jobs, doctor’s appointments, games, field trips**, and other regularly scheduled events will NOT be considered as a rare case. All religious holidays are recognized.

DO NOT ALLOW MAKE UP WORK GO UNTIL THE LAST MINUTE!

CLASSROOM MATERIALS:

You will need to obtain:

- 1) A CLASS NOTEBOOK (a loose-leaf notebook is best. You will end with about 500 pieces of paper).
- 2) A LAB NOTEBOOK
- 3) LOOSE LEAF PAPER
- 4) A CALCULATOR (GRAPHING OR SCIENTIFIC)
- 5) PENCILS, PENCILS AND MORE PENCILS!

CLASSROOM BEHAVIOR:

- 1) BE ON TIME (IN YOUR SEAT WHEN THE BELL RINGS)!
- 2) BE PREPARED (PENCIL, NOTEBOOK, CALCULATOR, ETC.)!
- 3) BE READY TO WORK!
- 4) BE COURTEOUS!

I - Basic Skills:

- 1) Differentiate between a solid, a liquid, and a gas at the macroscopic level.
- 2) Differentiate between a solid, a liquid, and a gas at the microscopic (particle) level.
- 3) Differentiate between a substance and a mixture; between an element and a compound; between a homogeneous mixture (a solution) and a heterogeneous mixture.
- 4) Distinguish between a physical and chemical property; between a physical and chemical change.
- 5) Describe the principle(s) between the following separation techniques and the kinds of mixtures to which they can be applied:
 - a) Filtration
 - b) Distillation
 - c) Crystallization
 - d) Chromatography
 - e) Extraction
- 6) Differentiate between the terms hypothesis, scientific law and scientific theory.
- 7) Identify the seven SI base units and the physical quantity they are used to describe.
- 8) Define the terms precision and accuracy as they relate to collected data.
- 9) Apply the rules of significant figures to calculations based upon experimental measurements.
- 10) Convert between Fahrenheit, Celsius, and Kelvin temperatures using the equations:
 - a) $F = 1.8C + 32$
 - b) $K = C + 273$
- 11) Use the equation: *density = mass/volume* to calculate one of these three quantities, given (or having calculated) the other two quantities.
- 12) Use dimensional analysis (the factor-label method) to convert lengths, volumes, masses, or other measures from one unit to another.

KNOW THESE TWO EQUATIONS BY HEART
II - Reading Schedule and Suggested Problems:

Day 0 [Aug. 2015] **Lab-1: Graphing in Microsoft Excel** [Due September 18th]

Day 1 [Tue. 9/8] **Read 1.1 to 1.3**
HW 1.1: 1.10, 1.12, 1.14, 1.18, 1.59, 1.65, 1.80

Lesson-1.1 **I** The Study of Matter
 II The Separation of Mixtures
 Lab-2: Accuracy and Precision

Day 2 [Wed. 9/9] **Read 1.4 to 1.6**
HW 1.2: 1.24bd&f, 1.34a&b, 1.38, 1.46, 1.75

Lesson-1.2 **III** Units and Measurement
 IV Experimental Error and Significant Figures
 V Conversion Factors and Dimensional Analysis
 VI Rules for Rounding

Review the material in ***Worksheet-1: Forms and Phases of Matter*** (pages 51-56) and work on the questions on page 56. This material is testable.

In AP Chemistry, you must know the names and the formulas (including the ionic charges) for these selected polyatomic ions **by heart**.

Name of the Polyatomic Ion	Chemical Formula	Name of the Polyatomic Ion	Chemical Formula
Ammonium	NH_4^{1+}	Mercury (I)	Hg_2^{2+}
Acetate	$\text{C}_2\text{H}_3\text{O}_2^{1-}$	Nitrite	NO_2^{1-}
Cyanide	CN^{1-}	Nitrate	NO_3^{1-}
Carbonate	CO_3^{2-}	Peroxide	O_2^{2-}
Hydrogen Carbonate (bicarbonate) old name	HCO_3^{1-}	Sulfite	SO_3^{2-}
Chromate	CrO_4^{2-}	Sulfate	SO_4^{2-}
Dichromate	$\text{Cr}_2\text{O}_7^{2-}$	Hydrogen Sulfate	HSO_4^{1-}
Permanganate	MnO_4^{1-}	Hydrogen Sulfite (bisulfite) old name	HSO_3^{1-}
Hypochlorite	ClO^{1-} or OCl^{1-}	Thiosulfate	$\text{S}_2\text{O}_3^{2-}$
Chlorite	ClO_2^{1-}	Phosphate	PO_4^{3-}
Chlorate	ClO_3^{1-}	Thiocyanate	SCN^{1-}
Perchlorate	ClO_4^{1-}	Hydronium	H_3O^{1+}
Oxalate	$\text{C}_2\text{O}_4^{2-}$	Hydroxide	OH^{1-}

The following is a table of selected metallic cations that have more than one possible charge.

You must know their traditional names, their IUPAC names, and their chemical formulas **by heart**.

Ion	Traditional Name	IUPAC Name*	Ion	Traditional Name	IUPAC Name*
Cu^{1+}	Cuprous	Copper (I)	Cu^{2+}	Cupric	Copper (II)
Fe^{2+}	Ferrous	Iron (II)	Fe^{3+}	Ferric	Iron (III)
Sn^{2+}	Stannous	Tin (II)	Sn^{4+}	Stannic	Tin (IV)
Pb^{2+}	Plumbous	Lead (II)	Pb^{4+}	Plumbic	Lead (IV)
Co^{2+}	Cobaltous	Cobalt (II)	Co^{3+}	Cobaltic	Cobalt (III)
Ni^{2+}	Nickelous	Nickel (II)	Ni^{3+}	Nickelic	Nickel (III)
Hg_2^{2+}	Mercurous**	Mercury (I)	Hg^{2+}	Mercuric	Mercury (II)

* IUPAC = International Union of Pure and Applied Chemistry

**Notice that the mercurous ion [mercury (I)] is a polyatomic cation.

Continues on page 6

I - Basic Skills:

- 1) Identify Dalton's four postulates of atomic theory.
- 2) Discuss the contributions of Thompson's Cathode-Ray Tube experiment, Millikan's Oil-Drop experiment, and Rutherford's Gold-Foil experiment to the evolution of atomic theory.
- 3) State the Laws of Conservation of Mass, Constant Composition, and Multiple Proportions and indicate how they provide evidence for the atomic theory.
- 4) Identify and give the relative charges and masses of the three major subatomic particles.
- 5) Relate a nuclear (nuclide) symbol to the number of protons and neutrons in the nucleus.
- 6) Given the masses and percent abundance of the isotopes of an element, calculate its atomic mass. Given the atomic mass of an element and the masses of its two isotopes, calculate their percent abundance.
- 7) Calculate the mass of an element in grams or in atomic mass units, given the other quantity.
- 8) Differentiate between an empirical formula and a molecular formula.
- 9) Classify a compound as ionic or molecular.
- 10) Given the formula of a species, determine the oxidation number of each atom.
- 11) Given either the name or the formula of an ionic compound, give the other.
- 12) Given either the name or the formula of an acid, give the other.
- 13) Given either the name or the formula of a molecular compound, give the other.
- 14) Given either the name or the formula of a simple alkane or alcohol, given the other.

II - Lecture Schedule and Assignments:

Day 3 [Thu. 9/10] **Read 2.1 to 2.4 [*Including "A Closer Look" on p.45 and p.48*]**
HW 2.1: 2.8, 2.10, 2.12, 2.18, 2.24, 2.30, 2.32

Lesson-2.1 I Atomic Theory
 II Atomic Structure
 III Atomic Weights

Day 4 [Fri. 9/11] **Read 2.5 to 2.7**
HW 2.2: 2.38, 2.46a&c, 2.48, 2.50, 2.54, 2.56

Lesson-2.2 IV The Periodic Table
 V Molecules and Molecular Compounds
 VI Ions and Ionic Compounds

Review the material about
ion found on pages 4 and 6.

Day 5 [Tue. 9/15] **Read 3.1 to 3.3**
HW 2.3: 2.64acd&f, 2.66ac&d, 2.68, 2.70, 2.72, 2.91 2.93, 2.96

Lesson-2.3 VII Naming Inorganic Compounds
 VIII Naming Organic Compounds

Lab-3: Density of an Unknown Solid

Review the material in **Worksheet-2: Measurement and the Mathematics of Chemistry** (pages 57-66) and work on the questions on pages 61, 63, 64, and 66. **This material is testable.**

Review the material in **Worksheet-3: Atoms, Isotopes, & Atomic Mass** (pages 67-74) and work on the questions on page 74. **This material is testable.**

The following is a table of selected metallic monoatomic cations. As shown on page 4, many metals can make cations with more than one charge. However, the metals shown below make cations with only one charge.

I call these metals the “*fabulous fifteen*” because there is only one charge. Know them **by heart**.

Ion	IUPAC Name	Ion	IUPAC Name	Ion	IUPAC Name
Li^{1+}	Lithium [†]	Fr^{1+}	Francium [†]	Ba^{2+}	Barium [†]
Na^{1+}	Sodium [†]	Be^{2+}	Beryllium [†]	Ra^{2+}	Radium [†]
K^{1+}	Potassium [†]	Mg^{2+}	Magnesium [†]	Ag^{1+}	Silver [†]
Rb^{1+}	Rubidium [†]	Ca^{2+}	Calcium [†]	Zn^{2+}	Zinc [†]
Cs^{1+}	Cesium [†]	Sr^{2+}	Strontium [†]	Al^{3+}	Aluminum [†]

[†]The charge of each of these ions is related to its position on the Periodic Table. **Figure it out.**

The following is a table of selected nonmetallic anions. You must know their IUPAC names as well as their chemical formulas **by heart**.

Ion	IUPAC Name	Ion	IUPAC Name	Ion	IUPAC Name
H^{1-}	Hydride ^{††}	O_2^{1-}	Superoxide	Cl^{1-}	Chloride ^{††}
N^{3-}	Nitride ^{††}	F^{1-}	Fluoride ^{††}	Se^{2-}	Selenide ^{††}
O^{2-}	Oxide ^{††}	P^{3-}	Phosphide ^{††}	Br^{1-}	Bromide ^{††}
O_2^{2-}	Peroxide	S^{2-}	Sulfide ^{††}	I^{1-}	Iodide ^{††}

^{††}The charge of each of these ions is related to its position on the Periodic Table. **Figure it out.**

Notes:

- 1) All polyatomic ions that end with the suffix “ate” and “ite” are negative. An “ate” polyatomic ion will always have one more oxygen atom than the “ite” polyatomic ion with the same root name. Thus you will know that phosphite is PO_3^{3-} .
- 2) All ions that end with the suffix “ide” are negative. All polyatomic ions ending with the suffix “ium” are positive.
- 3) The prefix “thio” indicates the replacement of an oxygen atom with a sulfur atom. Thus sulfate $[\text{SO}_4^{2-}]$ becomes thiosulfate $[\text{S}_2\text{O}_3^{2-}]$. The thiocyanate ion is SCN^{1-} , so the cyanate ion is OCN^{1-} .
- 4) Adding a hydrogen atom to a polyatomic ion lowers the negative charge by one. Thus you will know that hydrogen phosphate is HPO_4^{2-} and that dihydrogen phosphate is $\text{H}_2\text{PO}_4^{1-}$. The prefix “bi” in old traditional names indicates the presence of a hydrogen atom at the beginning of the ion.
- 5) If you know that the polyatomic ion ClO_3^{1-} is chlorate and if the chlorine atom is replaced with another atom from the same elemental family (the same group), then you will know that IO_3^{1-} is iodate and that BrO_3^{1-} is bromate. If PO_4^{3-} is phosphate, then AsO_4^{3-} is arsenate.
- 6) The traditional root names for some of the cations are based on the Latin names of the elements: copper [cuprum], iron [ferrum], tin [stannum], lead [plumbum], and mercury [mercurum].
 - ☞ Remember that “stann” is the old root name for tin. Think... “*Stan, Stan the tin man*”.
 - ☞ Remember that water pipes in homes were once made of lead and that the person who fixes water pipes is the plumber. So then remember that “plumb” is the root name for lead.
- 7) When naming cations, the “ic” ending always indicates the ion with the higher positive charge and the “ous” ending always indicates the ion with the lower positive charge.
- 8) The Roman numeral used in the modern IUPAC name is equal to the element’s oxidation number and is usually equal to the ion’s charge.

I - Basic Skills:

- 1) Be able to write and balance chemical equations given formulas or names.
- 2) Identify the chemical reactions: combination, decomposition, and combustion. [See page 10]
- 3) Given the formula of a species calculate its formula or molecular weight and its mole mass.
- 4) Relate the number of moles of a known formula to the number of grams or particles.
- 5) Given the formula of a compound, calculate the mass percents of the elements.
- 6) Determine the empirical formula of a compound given either the masses or the mass percents of the elements in the sample of the compound.
- 7) Determine the empirical formula of a compound given its mass and the masses of its combustion products.
- 8) Given the empirical formula of a compound and its molar mass, obtain its molecular formula.
- 9) Given the formula of an ionic compound, name the compound using either the Stock system (Roman numeral system) name or the traditional name. [See the summer assignment]
- 10) Given the name of an ionic compound [using either the Stock name or the traditional name] and a Periodic Table, write the formula for the ionic compound. This may include one that contains a transition metal ion and/or a polyatomic ion. [See the summer assignment]
- 11) Use a balanced equation to relate the number of moles or grams of reactants and products.
- 12) Given the number of moles or grams of each reactant, determine the limiting reactant and calculate the theoretical yield of product.
- 13) Relate the actual yield of product to the theoretical yield and percent yield.

II - Lessons and Assignments Schedule:Day 6 [Wed. 9/16] **Read** 3.1 to 3.4**HW 3.1:** 3.10, 3.14cd&e, 3.16, 3.24bd&f, 3.84b

Lesson-3.1

I

Chemical Equations: LO 1.17, LO 1.18, LO 3.1

II

Patterns of Chemical Reactions: LO 1.17, LO 1.18

III

Formula Weights and Percent Weights: LO 1.1, LO 1.3Day 7 [Thu. 9/17] **Read** 3.5 and 3.7**HW 3.2:** 3.36, 3.50a&b, 3.54

Lesson-3.2

IV

Avogadro's Number and the Mole: LO 1.4,

V

Empirical and Molecular Formulas: LO 1.2, LO 3.5, LO 3.6

It might be easier to change the milligrams to grams before starting the problems.

Day 8 [Fri. 9/18] **HW 3.3:** 3.52, 3.96a&b

Lesson-3.3

VI

Determining Formulas by Combustion Analysis: LO 1.2, LO 3.5, LO 3.6Day 9 [Mon. 9/21] **HW 3.4:** 3.8, 3.56, 3.58b, 3.74, 3.80

Lesson-3.4

VII

Calculations Based on Chemical Equations: LO 1.19, LO 3.3, LO 3.4Day 10 [Tue. 9/22] Student Activity: Determining an Element's Average Atomic Mass (p. 28) and Predicting Products and Writing Reactions (p. 8)**Lab-4:** Empirical Formula of a CompoundDay 11 [Thu. 9/24] **Review for TEST-1** (Chapters 1-3)Day 12 [Fri. 9/25] **TEST-1 (day-1) multiple choice:**Day 13 [Mon. 9/28] **TEST-1 (day-2) free response:**

(A) Combination Reactions: "two or more reactants (compounds and/or elements) form a compound".

- (1) **Two pure elements react ... to form a single compound.**
 - a) A clean strip of solid magnesium is burned in pure nitrogen gas.
 - b) Chlorine gas is passed over solid hot solid iron filings.
- (2) **A metallic oxide is added to water ... to form an ionic (hydroxide) base.**
 - a) Solid barium oxide is added to water.
 - b) Solid lithium oxide is added to water.
- (3) **A nonmetallic oxide is added to water ... to form an acid.**
 - a) Sulfur dioxide gas is bubbled into water.
 - b) Dinitrogen pentoxide gas is bubbled into water.
- (4) **Ammonia gas reacts with a hydrogen halide gas ... to form a solid ammonium salt.**
 - a) Ammonia gas and hydrogen chloride gas are mixed.
 - b) Ammonia gas and hydrogen bromide gas are mixed.
- (5) **Ammonia gas reacts with an aqueous acid ... to form an aqueous ammonium salt.**
 - a) Ammonia gas is bubbled into an aqueous solution of sulfuric acid.
 - b) Ammonia gas is bubbled into an aqueous solution of nitric acid.

(B) Decomposition Reactions: "A compound forms two or more products (compounds and/or elements)".

- (6) **Exposing hydrogen peroxide to a catalyst or bright light ... to form water & oxygen gas.**
 - a) Aqueous hydrogen peroxide is placed in sunlight.
 - b) Aqueous hydrogen peroxide is placed in the presence of the solid catalyst manganese dioxide.
- (7) **Heating a metallic carbonate ... to form a metallic oxide & carbon dioxide.**
 - a) Solid calcium carbonate is heated.
 - b) Solid sodium carbonate is heated.
- (8) **Heating a metallic hydrogen carbonate ... to form a metallic oxide, carbon dioxide & H₂O.**
 - a) Solid sodium hydrogen carbonate [Sodium bicarbonate] is heated.
- (9) **Heating ammonium carbonate ... to form ammonia, carbon dioxide & water.**
 - a) Solid ammonium carbonate is heated.

(C) Combustion Reactions: "A CH-compound or a CHO-compound plus O₂ form carbon dioxide and water".

- (10) **Burning a hydrocarbon in excess oxygen [or in air] ... to form carbon dioxide and water.**
 - a) Propane gas is burned in air.
- (11) **Burning an alcohol in excess oxygen [or in air] ... to form carbon dioxide and water.**
 - a) Liquid ethanol is burned in air.
- (12) **Burning a sugar in excess oxygen [or in air] ... to form carbon dioxide and water.**
 - a) Solid glucose is burned in air.

*****Reaction Categories A-C (Chapter-3)*****Reaction Categories D&E (Chapter-4)*****

(D) Double Replacement Reactions: "two compounds form two new compounds".

- (13) **Mixing an aqueous acid solution and an aqueous base solution ... to form a salt & water.**
 - a) An aqueous solution of hydrochloric acid is added to an aqueous solution of sodium hydroxide.
- (14) **Mixing two aqueous salt solutions ... to form a new aqueous salt solution & a precipitated salt.**
 - a) An aqueous solution of silver nitrate is added to an aqueous solution of sodium chloride.

(E) Single Replacement Reactions: "A compound and an element form a new compound and a new element".

- (15) **A metal is added to an aqueous acid solution ... to form a new salt & hydrogen gas.**
 - a) Solid zinc is added to an aqueous solution of acetic acid
- (16) **A metal is added to an aqueous salt solution ... to form a new salt & a metal.**
 - a) Solid aluminum is added to an aqueous solution of copper (II) nitrate.
- (17) **An alkali or alkaline earth metal is added to water ... to form a hydroxide base & hydrogen gas.**
 - a) Solid sodium is added to water.
- (18) **A halogen is added to an aqueous halogen salt solution ... to form a new halogen & a new halogen salt.**
 - a) Chlorine gas is bubbled into an aqueous solution of potassium iodide.

I - Basic Skills:

Identify the chemical reactions: single replacement (displacement) and double replacement and given the reactants' names, be able to predict the products and write a balanced molecular, complete ionic, or net ionic equation. [See page 10]

- 1) Identify any spectator ions in a solution reaction.
- 2) Be able to draw a set of particle diagrams for a single or double replacement reaction.
- 3) Classify a substance as an electrolyte [acid, base, salt, metallic oxide, or nonmetallic oxide] or a nonelectrolyte [sugar, alcohol, ketone, aldehyde, ether, or hydrocarbon]. [See page 12]
- 4) Classify an electrolyte as strong electrolyte or weak electrolyte.
- 5) Use the solubility guidelines to predict a precipitate in a solution reaction.
- 6) For redox reactions, identify the species being oxidized and the species being reduced.
- 7) Given the formula of a metal, predict the products it would produce in a salt solution, in an acid solution, or in pure water.
- 8) Use your knowledge of the Activity Series of Metals to determine whether a displacement (single replacement) reaction is spontaneous or nonspontaneous.
- 9) Given the formula of a species, determine the oxidation number for each element present.
- 10) Given or having calculated two of the three quantities: molarity, number of moles of solute, and number of liters of solvent, determine the other quantity.
- 11) Describe how to prepare a solution to a desired molarity, starting either with a pure solute or a concentrated solution of known molarity.
- 12) Given or having found a balanced equation (for a precipitation, redox, or acid-base reaction) as well as the volume of a known standard solution needed to reach the end-point when titrated against an unknown, determine the concentration, the number of moles, or the molar mass (given the mass) of the unknown.

II - Lessons and Assignments Schedule:

Day 14 [Tue. 9/29] **Read** 4.1 and 4.2

HW 4.1: 4.2, 4.4, 4.14, 4.18

Lesson-4.1 **I** General Properties of Aqueous Solutions: LO 2.8, LO 2.15, LO 3.1

Lab-5: Paper Chromatography

Day 15 [Wed. 9/30] **Read** 4.3 and 4.4

HW 4.2: 4.6, 4.10, 4.20, 4.24, 4.26

Lesson-4.2 **II** Precipitation Reactions

Day 16 [Thu. 10/1] **Read** 4.3 and 4.4

HW 4.3: 4.38, 4.40a&b, 4.42, 4.93b&e

Lesson-4.3 **III** Acid-Base Reactions: LO 1.17, LO 1.18, LO 2.1, LO 3.2, LO 3.4

Day 17 [Fri. 10/2] **Read** 4.5

HW 4.4: 4.46, 4.50, 4.52,

Lesson-4.4 **IV** Oxidation-Reduction Reactions: LO 1.17, LO 1.18, LO 3.4, LO 3.8

Day 18 [Mon. 10/5] **Read** 4.5

HW 4.5: 4.54, 4.56, 4.96

Lesson-4.5 **V** Displacement Reactions

Day 19 [Tue. 10/6] **Lab-6:** Quantitative Analysis of a Chemical Reaction

HW 4.6: 4.60, 4.64, 4.70, 4.74

Lesson-4.6 VI Solution Composition: LO 1.4, LO 2.9

Day 21[Thu. 10/8] Read 5.1 and 5.2

HW 4.7: 4.78, 4.80b, 4.84, 4.102, 4.106

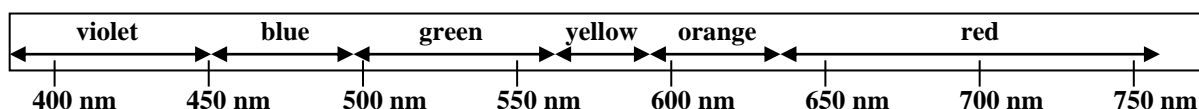
Lesson-4.7 VII Solution Stoichiometry-Titrations: LO 1.20, LO 3.9

Day 22[Fri. 10/9] Student Activity: Titration Problems ← DAY OF THE GO HOME EARLY DRILL

Ions that form Soluble Compounds	Exceptions	Ions that form Insoluble Compounds	Exceptions
Group-1 cations & NH_4^{1+}		CO_3^{2-}	Group-1 cations & NH_4^{1+}
NO_3^{1-}		S^{2-}	Group-1 cations & NH_4^{1+}
$\text{C}_2\text{H}_3\text{O}_2^{1-}$ [$\text{CH}_3\text{COO}^{1-}$]		PO_4^{3-}	Group-1 cations & NH_4^{1+}
HCO_3^{1-}		CrO_4^{2-}	Group-1 cations & NH_4^{1+} , & Mg^{2+} , Ca^{2+}
ClO_3^{1-} & ClO_4^{1-}		OH^{1-}	Group-1 cations & NH_4^{1+} , & Ca^{2+} , Sr^{2+} , Ba^{2+}
Cl^{1-} , Br^{1-} , I^{1-}	Ag^{1+} , Pb^{2+} , Hg_2^{2+}		
SO_4^{2-}	Ag^{1+} , Pb^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+}		

Compound	How to Identify the Classification of a Compound Based on Its Formula
Acid	<ul style="list-style-type: none"> Acid formulas usually start with hydrogen (H). Organic acid formulas either start with C's and H's and then end with the carboxyl group, (COOH) or start with H and end with $\text{C}_x\text{H}_y\text{O}_z$.
Base	<ul style="list-style-type: none"> Base formulas usually start with a metal or a (+) P.A.I. and end with the (-) P.A.I. hydroxide (OH^-). Organic base formulas can start with elements carbon and hydrogen and then end with the amine group (NH_2). Ammonia (NH_3) is a common base.
Salt	<ul style="list-style-type: none"> Salt formulas can start with a metal and end with a nonmetal (except for oxygen). Salt formulas can start with a (+) P.A.I. and end with a (-) P.A.I. (except for OH^-). Salt formulas can start with a metal and end with a (-) P.A.I. (except for OH^-). Salt formulas can start with a (+) P.A.I. and end with a nonmetal (except for oxygen). <i>Salts usually start with a metal or the ammonium ion.</i> <i>Simple salts start with a metal and end with a halogen.</i>
Binary Oxide	<ul style="list-style-type: none"> Metallic binary oxide formulas start with a metal and end with oxygen. Nonmetallic binary oxide formulas start with a nonmetal and end with oxygen.
Sugar	<ul style="list-style-type: none"> Sugar formulas will be composed of the three elements: carbon, hydrogen, and oxygen ($\text{C}_x\text{H}_y\text{O}_z$). The elements hydrogen and oxygen in a sugar will almost always be in a ratio of 2 to 1 ($\text{Y}:\text{Z} = 2:1$) but not always. <i>Example:</i> $\text{C}_5\text{H}_{10}\text{O}_5$ is ribose but $\text{C}_5\text{H}_{10}\text{O}_4$ is deoxyribose.
Alcohol	<ul style="list-style-type: none"> Alcohol formulas start with elements carbon and hydrogen and then end with the hydroxyl group – OH.
Ketone	<ul style="list-style-type: none"> Ketone formulas start with C's & H's, and then have one CO in the middle, and then ends with more C's & H's
Aldehyde	<ul style="list-style-type: none"> Aldehyde formulas start with C's & H's and then end with "CHO".
Ether	<ul style="list-style-type: none"> Ether formulas start with C's & H's, and then have one O, and then end with more C's & H's

Flame Test Colors for Some Alkali Metal Ions & Alkaline Earth Metal Ions						
Metal Ion	Na^{1+}	Li^{1+}	K^{1+}	Ca^{2+}	Sr^{2+}	Ba^{2+}
Color	Dark Yellow	Red (Crimson)	Violet (Lilac)	Orange-Red	Red-Scarlet	Yellowish Green



The basic colors in the visible spectrum change approximately every 50 nm.

I - Basic Skills:

- 1) Describe and calculate energy (the capacity to do work or to transfer heat) changes.
- 2) Know that **thermal energy** is the energy a substance possesses due to its temperature.
- 3) Relate the direction of heat flow in a reaction (exothermic or endothermic) to the sign of ΔH .
- 4) Use the First Law of Thermodynamics to calculate one of three quantities (ΔE , q , or w) knowing the values of the other two quantities.
- 5) Knowing the value of ($P\Delta V$) for a reaction, relate ΔH to ΔE ... [$\Delta H = \Delta E + P\Delta V$].
- 6) Given a thermochemical equation, calculate:
 - a) ΔH for a specific amount of reactant or product.
 - b) the amount of reactant or product required to produce a specified ΔH .
- 6) Use Hess' Law to calculate ΔH_R for
 - a) a stepwise reaction, knowing the value of ΔH_R for each step.
 - b) a step in a reaction, knowing the ΔH_R for every other step and for the overall reaction.
- 7) Use the general relation: $\Delta H_R = \sum \Delta H_f \text{ products} - \sum \Delta H_f \text{ reactants}$ to calculate:
 - a) ΔH_R for a reaction, knowing the heats of formation (ΔH_f) of all species.
 - b) ΔH_f for one species, knowing the ΔH_f of all other species as well as the ΔH_R .
- 8) Use the equations; $q = C\Delta T$, $q = mc\Delta T$, or $q = [mc+B]\Delta T$, to calculate one of the quantities, given (or having calculated) the other quantities.
- 9) Use calorimetric data (temperature change, mass of water, specific heat of water, calorimeter constant, bomb constant) to determine the heat flow for a reaction.

II - Lessons and Assignments Schedule:Day 23[Tue. 10/13] **Read** 5.3 and 5.4**HW 5.1:** 5.3, 5.20, 5.24, 5.26, 5.30

Lesson-5.1 I The Nature of Energy: LO 5.3
 II The First Law of Thermodynamics: LO 5.4

Lab-7: Qualitative Analysis of SolutionDay 24[Wed. 10/14] **Read** 5.5 and 5.6**HW 5.2:** 5.5, 5.32, 5.34, 5.38, 5.44

Lesson-5.2 III Enthalpy: LO 5.5, LO 5.6
 IV Enthalpy of Reaction:

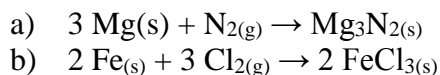
Day 25 [Thu. 10/15] **Read** 5.7 and 5.8**HW 5.3:** 5.48, 5.52, 5.54, 5.58, 5.62, 5.64

Lesson-5.3 V Calorimetry: LO 3.11, LO 5.3, LO 5.5, LO 5.6, LO 5.7
 VI Hess's Law:

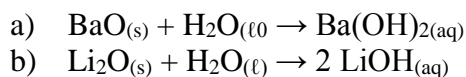
Day 26[Fri. 10/16] **Read** 6.1 and 6.2**HW 5.4:** 5.7, 5.8, 5.68a&c, 5.72b&d, 5.78Lesson-5.4 VII Enthalpy of Formation: LO 5.8

[Answers to questions from page 8]

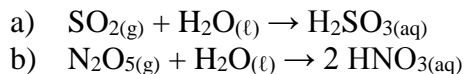
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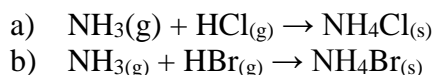
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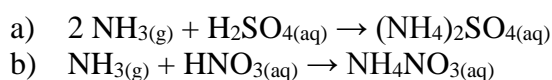
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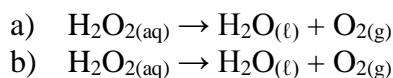
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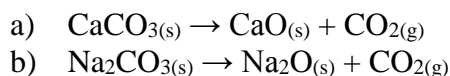
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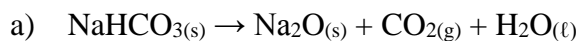
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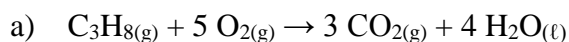
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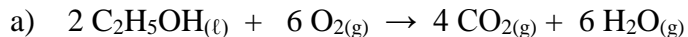
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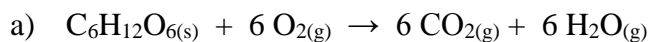
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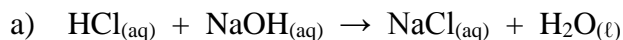


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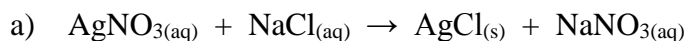


*******Reaction 1 – 12 are from Chapter-3*********Reaction 13 – 18 are from Chapter-4*******

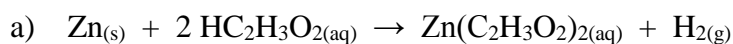
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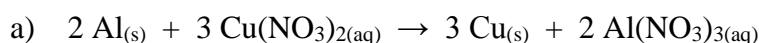
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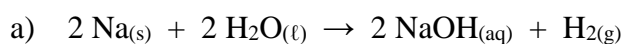
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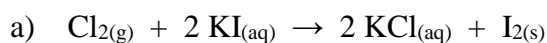
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I - Basic Skills:

- 1) Given a length in meters, convert to nanometers and vice versa.
- 2) Given one of the three quantities: λ , ν , ΔE , calculate the other two quantities.
- 3) Given a photon's wavelength in the visible spectrum, be able to predict its color [*See page 12*].
- 4) Be able to discuss the contributions of **Einstein, Planck, Bohr, Heisenberg, de Broglie, and Schrödinger** to the development of quantum theory.
- 5) Be able to explain Hund's rule, the Pauli exclusion principle, and the aufbau ordering principle.
- 6) Describe Planck's quantum hypothesis and explain how Einstein made use of it in his theory of the photoelectric effect.
- 7) Given either the wavelength of light that would decompose a chemical bond or the bond energy in kJ/mol, calculate the other.
- 8) Use the Bohr Theory to calculate the energy of an electron in a given principal energy level of the hydrogen atom, or the energy difference between two energy levels.
- 9) Given the energy absorbed/emitted by an electron moving between principal energy levels, determine either the electron's initial or final principal energy level given the other.
- 10) Given two of the three quantities for a particle: λ , ν , m , calculate the other quantity.
- 11) State the relationship between the wave function (ψ) at a given point in space and the probability of finding an electron at that point in space.
- 12) Describe the shapes of the s-orbitals, the p-orbitals, and the d-orbitals.
- 13) Given the atomic number of an element, write the electron configuration of its atom.
- 14) Given, or having derived, the electron configuration of an atom, draw its orbital structure.
- 16) Identify the various forms & uses of spectroscopy: emission, absorption, mass, photoelectron.
- 17) Be able to use photoelectron spectroscopy data to describe the electron structure of an atom.
- 18) Be able to use mass spectroscopy data to identify elements and the masses of specific atoms.

II - Lessons and Assignments Schedule:Day 27 [Mon. 10/19] **Read** 6.3 and 6.4**HW 6.1:** 6.22, 6.26, 6.76, 6.78, 6.82Lesson-6.1 I
II**Wave Nature of Light:**
Particle Nature of Light:Review the material in **Worksheet-4: the Electronic Structure of Atoms** (pages 75-84). This material is testable.Day 28 [Tue. 10/20] **Read** 6.5 and 6.6**HW 6.2:** 6.32, 6.34a&b, 6.36b&c, 6.38a, 6.100

Lesson-6.2 III

The Bohr Model of the Atom: LO 1.12**Lab-8:** Heat of NeutralizationDay 29 [Wed. 10/21] **Read** 6.7 to 6.9**HW 6.3:** 6.7, 6.42, 6.44, 6.48c, 6.54, 6.56, 6.89

Lesson-6.3 VI

The Quantum Mechanical Model of the Atom: LO 1.12

Atomic radii vary from about 30 pm to 300 pm.

Element 118 currently uses the temporary symbol "Uuo"

Day 30 [Thu. 10/22] **HW 6.4:** 6.60, 6.62a, 6.64b&d, 6.68bd&f, 6.74, 6.98

Lesson-6.4 V

Many-Electron Atoms and Electron Structures: LO 1.9, LO 1.13Day 31 [Fri. 10/23] **HW 6.5:** Homework WorksheetLesson-6.5 VI **Spectroscopy:** LO 1.6, LO 1.7, LO 1.14, LO 1.15, LO 1.16Day 32 [Mon. 10/26] **Read** 7.1 to 7.3**HW 7.1:** 7.8, 7.10b, 7.12, 7.24, 7.28, 7.30a&d, 7.34Lesson-7.1 I
II**Periodic Table:**
Sizes of Atoms and Ions: LO 1.5, LO 1.8

Day 33 [Tue. 10/27] **Lab-9: Law of Dulong and Petit**

Day 34 [Wed. 10/28] **Review for TEST-2 (Chapters 4-6)**

Day 35 [Thu. 10/29] **TEST-2 (day-1) multiple choice:**

Day 36 [Fri. 10/30] **TEST-2 (day-2) free response:**

Student Activity-2: Reactions of Metals, Nonmetals, and their Oxides: Write net ionic equations.

1) Write a balanced equation for the reaction between pure barium and pure iodine.

2) Write a balanced equation for the reaction between pure calcium and water.

3) Write a balanced equation for the reaction between pure sodium and pure oxygen.

4) Write a balanced equation for the reaction between carbon dioxide gas and water.

5) Write a balanced equation for the reaction between pure strontium oxide and aqueous hydrochloric acid.

6) Write a balanced equation for the reaction between dinitrogen pentoxide and aqueous lithium hydroxide.

7) Write a balanced equation for the reaction between pure potassium and pure oxygen.

8) Write a balanced equation for the reaction between pure calcium oxide and water.

9) Write a balanced equation for the reaction between pure lithium and pure oxygen.

10) Write a balanced equation for the reaction between zinc and aqueous sulfuric acid.

11) Write a balanced equation for the reaction between pure sulfur and pure oxygen.

12) Write a balanced equation for the reaction between pure white phosphorus and excess pure oxygen.

I - Basic Skills:

- 1) Be able to discuss the contributions of Dalton, Berzelius, Dobereiner, Newlands, Mendeleev, and Moseley to the development of the modern Periodic Table of Elements.
- 2) Determine the value of the ideal effective nuclear charge or the value of the ideal shielding constant experienced by an electron in an atom and explain the difference from their actual values in terms of electron repulsions, electron shielding, and/or electron penetration.
- 3) Explain the trends in the Periodic Table with respect to ionization energy, electron affinity, atomic radius, metallic character, and nonmetallic character in terms of effective nuclear charge, electron repulsions, electron shielding, and electron penetration.
- 4) State the electron configuration of monoatomic ions, including those of transition elements.
- 5) Compare the radii of monoatomic ions to their neutral atoms and explain the differences in terms of effective nuclear charge, electron repulsions, and/or electron shielding.
- 6) Describe, with the aid of balanced equations, the reactions of metals with nonmetals, of metals with water, and of metals with an acid. **Given reactants be able to predict products**
- 7) Describe, with the aid of balanced equations, the reactions of metals with oxygen to produce oxides, peroxides, or superoxides. **Given reactants be able to predict products**
- 8) Describe, with the aid of balanced equations, the reactions of metallic oxides with water and of nonmetallic oxides with water. **Given reactants be able to predict products**
- 9) Describe, with the aid of balanced equations, the reactions of metallic oxides with acids and of nonmetallic oxides with bases. **Given reactants be able to predict products**
- 10) Describe several chemical and physical properties of active metals [alkali metals & large alkaline earth metals], and active nonmetals [halogens].

II - Lessons and Assignments Schedule:Day 32 [Mon. 10/26] **Read 7.1 to 7.3****HW 7.1:** 7.8, 7.10b, 7.12, 7.24, 7.28, 7.30a&d, 7.34

Lesson-7.1 I Periodic Table:
 II Sizes of Atoms and Ions: LO 1.5, LO 1.8

Day 33 [Tue. 10/27] **Lab-9: Law of Dulong and Petit**Day 34 [Wed. 10/28] **Review for TEST-2 (Chapters 4-6)**Day 35 [Thu. 10/29] **TEST-2 (day-1) multiple choice:**Day 36 [Fri. 10/30] **TEST-2 (day-2) free response:**Day 37 [Mon. 11/2] **Read 7.6 to 7.8****HW 7.2:** 7.38a&c, 7.42a&b, 7.44ac&e, 7.46, 7.50, 7.97a&c

Lesson-7.2 III Ionization Energy: LO 1.5, LO 1.6, LO 1.7
 IV Electron Affinity: LO 1.5, LO 1.6

Day 38 [Tue. 11/3] **HALF-DAY:****Student Activity:** Reactions of Metals, Nonmetals, and Their Oxides (p. 16 & 14)Day 39 [Wed. 11/4] **Read 8.1 and 8.2****HW 7.3:** 7.60, 7.64, 7.68, 7.70, 7.76, 7.110a&c

Lesson-7.3 V Properties of Metals, Nonmetals, and Metalloids: LO 1.9, LO 1.10
 VI Group Trends in Metals and Nonmetals: LO 1.9, LO 1.10

Reactions of Metals, Nonmetals, and Their Oxides

- 1) Pure Metal + Pure Nonmetal → Solid Ionic Compounds
- 2) Pure Nonmetal + Pure Nonmetal → Gaseous Molecular Compounds
- 3) Pure Metal* + Pure Oxygen gas → Solid Metallic Oxide
 - * **notable exception**: Pure Sodium + Pure Oxygen gas → Sodium Peroxide
 - * **notable exception**: Pure Potassium + Pure Oxygen gas → Potassium Superoxide
 - * **notable exception**: Pure Cesium + Pure Oxygen gas → Cesium Superoxide
- 4) Pure Nonmetal + Pure Oxygen gas → Gaseous Nonmetallic Oxide
- 5) Pure Metal* + Acid → Salt (Aqueous or Solid) + Hydrogen Gas
 - * **notable exceptions**: Au, Ag, & Cu. *These metals tend NOT to react with most acids.*
 Au will react with a 3:1 mixture of HCl & HNO₃ known as ***aqua regia*** (royal water).
 Ag and Cu will react with HNO₃ but do NOT produce hydrogen gas.
- 6) Pure Alkali Metal + Water → Aqueous Base + Hydrogen Gas
 Pure Lower Alkaline Earth Metal + Water → Aqueous Base + Hydrogen Gas
- 7) Metallic Oxide + Water → Base (aqueous or solid)
 Metallic Peroxide + Water → hydrogen Peroxide + Base (aqueous or solid)
 Metallic Superoxide + Water → hydrogen Peroxide + Base (aqueous or solid) .+ Oxygen Gas
- 8) Nonmetallic Oxide + Water → Aqueous Acid
- 9) Metallic Oxide + Acid → Salt (aqueous or solid) + Water
- 10) Nonmetallic Oxide + Base → Salt (aqueous or solid) + Water

Important Alkali and Alkaline Earth Compounds

- 1) Important Lithium Compound:
 Li₂CO₃ - “lithium” [*treatment for manic-depression*]
- 2) Important Sodium Compounds:
 NaCl - **table salt, rock salt** (mineral name is **halite**)
 NaOH - **lye** [*used as a drain cleaner and in making soap*]
 Na₂CO₃•10H₂O - washing soda [*a laundry detergent booster*]
 NaHCO₃ - **baking soda** (bicarbonate of soda)
 NaClO - **household bleach** (in water)
- 3) Important Potassium Compounds:
 KNO₃ - **saltpeter** [*used in making gunpowder and fertilizer*]
 K₂CO₃ (KOH) - potash (from wood ashes) [*used in making glass or soap*]
- 4) Important Magnesium Compounds:
 MgSO₄•X H₂O - **Epsom salts** [*used as a laxative or as bath salts*]
 Mg(OH)₂ - **milk of magnesia** when mixed with water [*used as an antacid*]
- 5) Important Calcium Compounds:
 CaCO₃ - **limestone** (mineral name is **calcite**) [*used as chalk*]
 CaO - **lime** [*used in making cement and to “sweeten” (de-acidify) lawns*]
 CaSO₄•2H₂O - gypsum [*used in making wall board*]
 CaSO₄•½H₂O - plaster of Paris [*used in making casts*]

You must memorize the formulas of the names boldfaced

I - Basic Skills:

- 1) Having located a main group (groups 1, 2, 13, 14, 15, 16, 17, & 18) element in the Periodic Table, state its outer (valence) electron configuration.
- 2) Write the Lewis models (Lewis structures) for neutral atoms and for monoatomic ions.
- 3) Write Lewis models for binary ionic compounds, covalent compounds (molecules), polyatomic ions, and compounds containing polyatomic ion(s).
- 4) Use formal charge to evaluate different Lewis models for the same molecule or ion.
- 5) Predict whether a species will show resonance, given or having derived its Lewis model.
- 6) Using only the Periodic Table, describe the electronegativity trends of the elements and compare the relative polarity of different bonds.
- 7) Compare bond distances and bond enthalpies for polar bonds versus nonpolar bonds as well as multiple bonds versus single bonds.
- 8) Use a table of bond enthalpies to calculate ΔH for gas phase reactions.
- 9) Discuss the lattice energy trends among ionic compounds (ie: LiCl, NaCl, and KCl, or NaF, NaCl, and NaBr, or KF, CaO, and ScN) and how they relate to trends in melting point or solubility.

II - Lessons and Assignments Schedule:Day 40 [Thu. 11/5] **Read 8.3 and 8.4****HW 8.1:** 8.1, 8.3, 8.8, 8.12, 8.16, 8.20

Lesson-8.1

I

Chemical Bonding:

II

Ionic Bonding: LO 1.8, LO 1.11, LO 2.17

Write the chemical formulas

Day 41 [Fri. 11/6]

Read 8.5 and 8.6**HW 8.2:** 8.36a&b, 8.40a&b, 8.42, 8.44, 8.79

Lesson-8.2

III

Covalent Bonding: LO 2.17, LO 2.18, LO 5.1

Day 42 [Mon. 11/9]

Read 8.7**HW 8.3:** 8.46ab&c, 8.52, 8.62

Lesson-8.3

IV

Lewis Models and Resonance Structures:

V

Exceptions to the Octet Rule:

Day 43 [Tue. 11/10]

Lab-10: Precipitation Titration

Day 44 [Thu. 11/12]

Read 8.8**HW 8.4:** 8.48, 8.50, 8.64, 8.86

Lesson-8.4

VI

Oxidation Number and Formal Charge:

Day 45 [Fri. 11/13]

Read 9.1 to 9.2**HW 8.5:** 8.66a&c, 8.70, 8.88a&d, 8.99

Lesson-8.5

VII

Bond Enthalpy: LO 5.8

Student Activity: Drawing a Lewis model for an Unknown Substance

10.000 g of an unknown solid white substance was analyzed and found to be composed of carbon, hydrogen, and oxygen. Its molar mass was determined to be 126.08 g/mole. Its melting point was found to be between 101°C and 102°C. When heated gently over a period of time, it released a vapor. When some cobalt chloride indicator paper was held in the vapor, it turned from blue to pink. After having been cooled, it was found to have lost 2.859 g of its mass. 1.000 g of the cooled sample was analyzed and found to have a molar mass of 90.04 g/mol. A second 1.000 g cooled sample was found to dissolve in water. Its aqueous solution had a pH less than 7 and it conducted an electric current. The remaining cooled sample was burned and its combustion products were 5.026 g of carbon dioxide and 1.029 g of water. Based on this data, (a) explain why it lost mass when heated gently. Of the sample that was burned, determine both its (b) empirical formula $[H_xC_yO_z]$ and (c) molecular formula. (d) Determine the molecular formula of the original sample before it was heated gently. (e) Draw a Lewis model for the sample that was burned that conforms to all of this data (all carbons are bonded to each other, each carbon has an equal number of oxygen atoms bonded to it, and each hydrogen atom is bonded to an oxygen atom).

I - Basic Skills:

- 1) In terms of electron pairs, state how a molecule's geometry is predicted by VSEPR theory.
- 2) Given or having derived the Lewis model for a molecule or polyatomic ion, apply VSEPR principles to predict (i) the number of electron domains around the central atom, (ii) its electron domain geometry, (iii) its molecular geometry, (iv) the bonding angle(s) present, (v) the hybridization of the central atom, (vi) the number of sigma (σ) and pi (π) bonds present and (vii) whether it will have a net dipole moment (μ) equal to or greater than zero.
- 3) Draw the hybridized orbital model for the central atom of a molecule or polyatomic ion.
- 4) Explain how pi-bonding produces a rigid structure that results in cis and trans isomers.

II - Lessons and Assignments Schedule:Day 46 [Mon. 11/16] **Read 9.3 to 9.4****HW 9.1:** 9.1, 9.2, 9.12a, 9.16ab&c, 9.20b&c, 9.22b&c, 9.26a&d 9.75Lesson-9.1 I Molecular Shapes: VSEPR Theory:Day 47 [Tue. 11/17] **Read 9.5****HW 9.2:** 9.34, 9.36[Draw the Lewis structure for each compound], 9.38, 9.81Lesson-9.2 II Polarity of Polyatomic Molecules: LO 20.21**Lab-11:** Activity Series of Selected MetalsDay 48 [Wed. 11/18] **Read 9.6****HW 9.3:** 9.5, 9.7, 9.40, 9.44, 9.48Lesson-9.3 III Valance Bond (VB) Theory: LO 20.21Day 49 [Thu. 11/19] **HW 9.4:** 9.50, 9.52^{**}, 9.56a&b, 9.58Lesson-9.4 VI Multiple Bonds and the Valance Bond Theory:

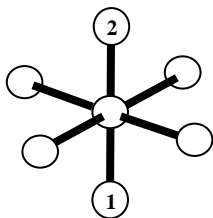
[**Draw the Lewis structures for N₂ and N₂H₄ to support your answer in 9.52]

Day 50 [Fri. 11/20] **HW 9.5:** 9.62, 9.66a^{**}, 9.68a&b, 9.70[**Calculate the MO bond order for O₂²⁻ & O₂¹⁻ to support your answer in 9.66a]Lesson-9.5 V Molecular Orbitals:Day 51 [Mon. 11/23] **Read 10.1 to 10.3****HW 10.1:** 10.2, 10.14, 10.18ab&c, 10.20a, 10.22, 10.24, 10.26, 10.89Lesson-10.1 I Characteristics of Gases: Pressure: LO 2.4II The Gas Laws: LO 2.5, LO 2.6Day 52 [Tue. 11/24] Student Activity: Drawing a Lewis Model of an Unknown Substance (p. 18)**Read 10.4 and 10.5****HW 10.2:** 10.3a, 10.32, 10.36, 10.46, 10.48, 10.52 ← **Due Tuesday 12/1**Lesson-10.2 III The Ideal-Gas Equation: LO 2.5, LO 2.6IV Further Applications of the Ideal-Gas Equation: LO 2.5, LO 2.6, LO 3.4**HAPPY THANKSGIVING!**Day 53 [Mon. 11/30] **Review for TEST-3** (Chapters 7-9)Day 54 [Tue. 12/1] **Lab-12:** Molar Mass of an Unknown Volatile GasDay 55 [Wed. 12/2] **TEST-3 (day-1) multiple choice:**Day 56 [Thu. 12/3] **TEST-3 (day-2) free response:**

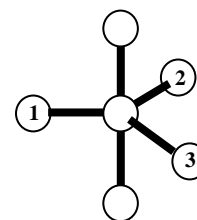
VSEPR Theory: Octet Based Geometries				VB Theory: Hybridization	
# of sigma bonds on central atom	# of lone pairs on central atom	Electron Domain Geometry	Molecular Geometry	hybridization present	hybridization of the s and p orbitals
2	0	linear	linear	sp	$\square \square \square \rightarrow \square \square$ s p sp p
2	1	trigonal planar	bent	sp ²	$\square \square \square \rightarrow \square \square \square$ s p sp ² p
3	0	trigonal planar	trigonal planar	sp ²	$\square \square \square \rightarrow \square \square \square$ s p sp ² p
2	2	tetrahedral	bent	sp ³	$\square \square \square \rightarrow \square \square \square$ s p sp ³
3	1	tetrahedral	trigonal pyramidal	sp ³	$\square \square \square \rightarrow \square \square \square$ s p sp ³
4	0	tetrahedral	tetrahedral	sp ³	$\square \square \square \rightarrow \square \square \square$ s p sp ³

VSEPR Theory: Expanded Octet Based Geometries				VB Theory: Hybridization *	
# of sigma bonds on central atom	# of lone pairs on central atom	Electron Domain Geometry	Molecular Geometry	hybridization present	hybridization of the s, p, and d orbitals
2	3	trigonal bipyramidal	linear	sp ³ d	$\square \square \square \square \rightarrow \square \square \square$ d → d $\square \square \square$ sp ³ d s p
3	2	trigonal bipyramidal	T-shaped	sp ³ d	$\square \square \square \square \rightarrow \square \square \square$ d → d $\square \square \square$ sp ³ d s p
4	1	trigonal bipyramidal	see saw	sp ³ d	$\square \square \square \square \rightarrow \square \square \square$ d → d $\square \square \square$ sp ³ d s p
5	0	trigonal bipyramidal	trigonal bipyramidal	sp ³ d	$\square \square \square \square \rightarrow \square \square \square$ d → d $\square \square \square$ sp ³ d s p
4	2	octahedral	square planar	sp ³ d ²	$\square \square \square \square \rightarrow \square \square \square$ d → d $\square \square \square$ sp ³ d ² s p
5	1	octahedral	square pyramidal	sp ³ d ²	$\square \square \square \square \rightarrow \square \square \square$ d → d $\square \square \square$ sp ³ d ² s p
6	0	octahedral	octahedral	sp ³ d ²	$\square \square \square \square \rightarrow \square \square \square$ d → d $\square \square \square$ sp ³ d ² s p

*There is much controversy about whether the d-orbitals play such a role in bonding as this model suggests. However, until such time that this issue is resolved, this model will still be presented in its current form.



The numbers in the two models show the sequence of how bonded atoms are replaced by non-bonded electron pairs.



I - Basic Skills:

- 1) Use the Ideal Gas Law to:
 - a) determine the effect of a change in one or more of the variables (P, V, n, or T) on the value of one of these variables.
 - b) solve for one variable (P, V, n, or T), knowing the values of the other three variables.
 - c) calculate the density of a gas, given the necessary data (molar mass, pressure, and temperature).
 - d) calculate the molar mass of a gas, given the necessary data (density, pressure, and temperature).
 - e) relate the volume of a gas involved in a reaction to the amount (mass, volume, or number of particles) of another reactant or product.
- 2) Relate the volumes of gases (at the same temperature and pressure) in a reaction.
- 3) Use Dalton's Law to find the partial pressure of a gas in a mixture; relate partial pressure to mole fraction.
- 4) Use Graham's Law to relate the molar masses of two gases to their rates or times of effusion.
- 5) Use the van der Waal's equation to explain why real gases deviate from ideal gas behavior.

II - Lessons and Assignments Schedule:Day 51 [Mon. 11/23] **Read** 10.1 to 10.3**HW 10.1:** 10.2, 10.14, 10.18ab&c, 10.20a, 10.22, 10.24, 10.26, 10.89

Lesson-10.1 I Characteristics of Gases: Pressure: LO 2.4
 II The Gas Laws: LO 2.5, LO 2.6

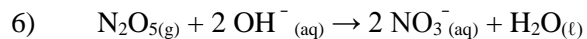
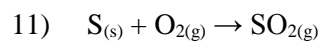
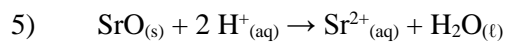
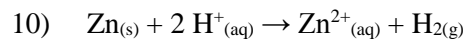
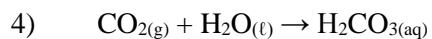
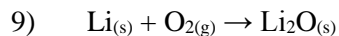
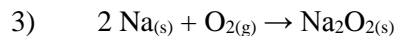
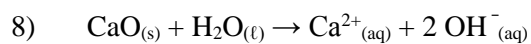
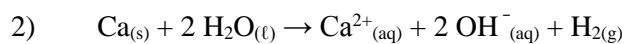
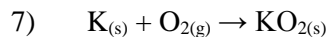
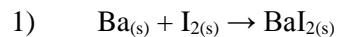
Day 52 [Tue. 11/24] Student Activity: Drawing a Lewis Model of an Unknown Substance (p. 18)**Read** 10.4 and 10.5**HW 10.2:** 10.3a, 10.32, 10.36, 10.46, 10.48, 10.52 ← **Due Tuesday 12/1**

Lesson-10.2 III The Ideal-Gas Equation: LO 2.5, LO 2.6
 IV Further Applications of the Ideal-Gas Equation: LO 2.5, LO 2.6, LO 3.4

HAPPY THANKSGIVING!Day 53 [Mon. 11/30] **Review for TEST-3** (Chapters 7-9)Day 54 [Tue. 12/1] **Lab-12:** Molar Mass of an Unknown Volatile GasDay 55 [Wed. 12/2] **TEST-3 (day-1) multiple choice:**Day 56 [Thu. 12/3] **TEST-3 (day-2) free response:**Day 57 [Fri. 12/4] **HW 10.3:** 10.5, 10.8, 10.56, 10.58, 10.66

Lesson-10.3 V Gas Mixtures and Partial Pressures: LO 2.4
 VI Kinetic Molecular Theory of Gases: LO 2.4, LO 5.2

Day 58 [Mon. 12/7] **Read** 11.1 and 11.3**HW 10.4:** 10.70, 10.74, 10.76, 10.84, 10.103Lesson-10.4 VII The Nonideal Behavior of Real Gases: LO 2.4, LO 2.12

[Answers to student activity from page 16]

I - Basic Skills:

- 1) Compare different molecular substances with respect to physical properties (melting point, boiling point, conductivity, water solubility, etc.) and types of intermolecular forces.
- 2) Use the Clausius-Clapeyron equation [$\ln(P_2/P_1) = -\Delta H_{\text{vap}}/R (1/T_2 - 1/T_1)$] to calculate one of the three quantities, P_2 , T_2 , or H_{vap} , given the values of the other two and P_1 and T_1 .
- 3) Draw a phase diagram for a pure substance, given appropriate data, and (1) state what phases are present at any given point or (2) what phase change occur between points on the diagram.
- 4) Distinguish among ionic, molecular, network covalent and metallic solids with regards to both particle structure, bonding and physical properties (melting point, boiling point, conductivity, water solubility, etc.).
- 5) Discuss the two models of metallic bonding: electron-sea model and molecular orbital model.
- 6) Describe the composition and properties of the following types of solid materials: ceramics, glass, metallic alloys, semiconductors, and plastics.
- 7) Differentiate between substitutional alloys and interstitial alloys.
- 8) Describe how the process of doping leads to changes in the properties of a semiconductor.
- 9) Differentiate between a n-type and a p-type semiconductor and explain how they both form the basis for diodes and transistors (the p-n junction).

II - Lessons and Assignments Schedule:Day 59 [Tue. 12/8] **Read 11.4****HW 11.1:** 11.2, 11.14, 11.16, 11.22, 11.30a&c, 11.32Lesson-11.1 **I** Kinetic Molecular Theory: LO 2.4**II** Intermolecular (van der Waal's) Forces: LO 2.3, LO 2.11, LO 2.13, LO 5.11**III** Viscosity and Surface Tension: LO 2.3, LO 2.13, LO 2.16**Lab-13:** Activity Series of the HalogensDay 60 [Wed. 12/9] **Read 11.5****HW 11.2:** 11.34, 11.38, 11.40 ← *convert 27.49 kJ/mol to J/g*Lesson-11.2 **IV** Changes of State: LO 2.3, LO 2.16, LO 5.6, LO 5.10Day 61 [Thu. 12/10] **Read 11.6 and 11.8****HW 11.3:** 11.46, 11.50a ..., 11.90b&c

... and then use the Clausius-Clapeyron equation to verify your prediction.

Lesson-11.3 **V** Vapor Pressure and Boiling Point:Day 62 [Fri. 12/11] **Read 11.7 (BE SURE TO READ SECTION 11.8)****HW 11.4:** Homework WorksheetLesson-11.4 **VI** Phase Diagrams:**VII** Classes of Solids: LO 2.3, LO 2.19, LO 20.20, LO 20.22Day 63 [Mon. 12/14] **Read 12.1-12.2, 22.10, 23.5-23.6****HW 11.5:** Homework WorksheetLesson-11.5 **VIII** Types of Solid Materials: LO 2.3

A General Comparison of the Properties of the Four Classes of Solids				
Physical or Chemical Property	Class of Solid			
	Ionic Solids	Metallic Solids	Covalent Solids	
			Network**	Molecular
Type of Chemical Bonding Present	Ionic or both ionic and covalent	Metallic, only	Covalent, only (between atoms)	Covalent, only (within polyatomic molecules)
Basic Particle(s) "Building Blocks"	Cations and anions	Cations, only	Individual atoms	Individual Molecules
Attractive Forces Broken during Melting & Boiling	Ionic Bonds	Metallic Bonds	Covalent Bonds	Van der Waal's (Intermolecular) Forces
Melting & Boiling Point	Very high	Low to very high	Very high	Low to very low
Volatility	Very low	Very low	Very low	High to very high
Electrical Conductivity	<u>Poor when solid</u> , but good when liquid or aqueous	Good (in all phases)	Usually poor	Usually poor
Thermal Conductivity	Usually poor	Good	Usually poor	Usually poor
Hardness ("scratchability")	Very hard	Medium to very hard	Very hard	Soft to very soft
Malleability/ Brittleness	Brittle	Malleable	Brittle	Brittle
Examples	NaCl, CaF ₂ , CuSO ₄ , MgO, ZnS, NH ₄ Cl, KClO ₃ , Ba(OH) ₂	Cu, Ag, Au, Fe, Al, brass, bronze	Si, SiC, Diamond (C), graphite (C), quartz (SiO ₂), BN, asbestos	H ₂ O, CO ₂ , Ar, I ₂ , C ₆ H ₂ O ₆ , O ₂ , CH ₃ OH, C ₂₅ H ₅₂

**The examples listed for the network solids are ones that you should know. (MEMORIZE)

I - Basic Skills:

- 1) Write balanced equations (or particle diagrams) to represent the dissolution of a solute (an electrolyte or a nonelectrolyte) in a solvent.
- 2) Discuss the energy changes that occur during the solution process.
- 3) Predict the effect of change in temperature or pressure upon the solubilities of solids and gases in solution.
- 4) Given appropriate data, calculate the molarity of a solution.
- 5) Relate the molarity of an electrolyte solution to the molarities of its ions.

II - Lessons and Assignments Schedule:

Day 64 [Tue. 12/15] **Read** 3.1 to 3.3 ↖ explain in terms of solute-solvent interactions
HW 13.1: 13.3, 13.4, 13.12-a-b&c, 13.20(show work), 13.22(show work)

Lesson-13.1 **I** General Properties of Solutions (Reviewed): LO 2.8, LO 2.15
 II The Solution Process: LO 2.8, LO 2.15
 III Factors Affecting Solubility: LO 2.8, LO 2.14, LO 5.10

Lab-14: Heat of Solution

Day 65 [Wed. 12/16] **Read** 13.4
HW 13.2: 13.5, 13.7, 13.34b, 13.42, 13.48a, 13.50

Lesson-13.2 **IV** Expressing Solution Concentration Qualitatively: LO 2.9
 V Expressing Solution Concentration Quantitatively: LO 2.9

Day 66 [Thu. 12/17] **1st Semester Exam Review** [See page 26 for suggested review]

Day 67 [Fri. 12/18] **Review for TEST-4** (Chapters 10-13)

Day 68 [Mon. 12/21] **TEST-4 (day-1) multiple choice:**

Day 69 [Tue. 12/22] **TEST-4 (day-2) free response:**

1st Semester Exam Review [See page 26 for suggested review]

Day 70 [Wed. 12/23] **1st Semester Exam Review** [See page 26 for suggested review]

THE END OF 1st SEMESTER MATERIAL!

HAPPY DECEMBER BREAK!

You may want to review for the 1st Semester Exam. See the suggested review on page 26.

Day 71 [Mon. 1/4] **1st Semester Exam Review** [See page 26 for suggested review]

Day 72 [Tue. 1/5] **1st SEMESTER EXAM (Part I:)**

1st Semester Exam Review [See page 26 for suggested review]

Day 73 [Wed. 1/6] **1st SEMESTER EXAM (Part II:)**

Day 74 [Thu. 1/7] **1st SEMESTER EXAM (Part III:)**

Day 75 [Fri. 1/8] **1st SEMESTER EXAM (Part IV:)**

Your first semester exam will take place during class in early January.

[Do not be absent! Do not be absent! Do not be absent! Do not be absent! Do not be absent!]

Day-1, Tuesday, January 5th: Chapters 1, 2, & 7

Review Questions from Chapter 1:

33 & 60

Review Questions from Chapter 2:

3, 17, 19, 59bdfhj, 61bdf, 65bcde

Review Questions from Chapter 7:

11, 21, 31

Day-2, Wednesday, January 6th: Chapters 6, 8, & 9

Review Questions from Chapter 6:

8, 17, 29, 39, 51, 67bf, 71ac

Review Questions from Chapter 8:

5, 19, 21, 43, 65a, 78, 84bd

Review Questions from Chapter 9:

6, 15, 17, 19, 29, 39, 49, 76

Day-3, Thursday, January 7th: Chapter 5, 10, & 11

Review Questions from Chapter 5:

4, 25, 37, 51, 55, 61, 73

Review Questions from Chapter 10:

33, 47, 59, 79, 80, 81

Review Questions from Chapter 11:

4, 13, 17, 55, 77, 89 [For 11.89, instead of plotting the data, use the Clausius-Clapeyron equation to determine the ΔH_{vap} and then use that value to determine the normal boiling point]

Day-4, Friday, January 8th: Chapters 3, 4, & 13

Review Questions from Chapter 3:

4, 23b, 43a, 45a, 47a, 51b, 59ab, 73a

Review Questions from Chapter 4:

1, 3, 49

Review Questions from Chapter 13:

1, 13, 41, the practice exercise at the bottom of page 556

Day-5, Tuesday, January 12th: The Return and Review of the First Semester Exam

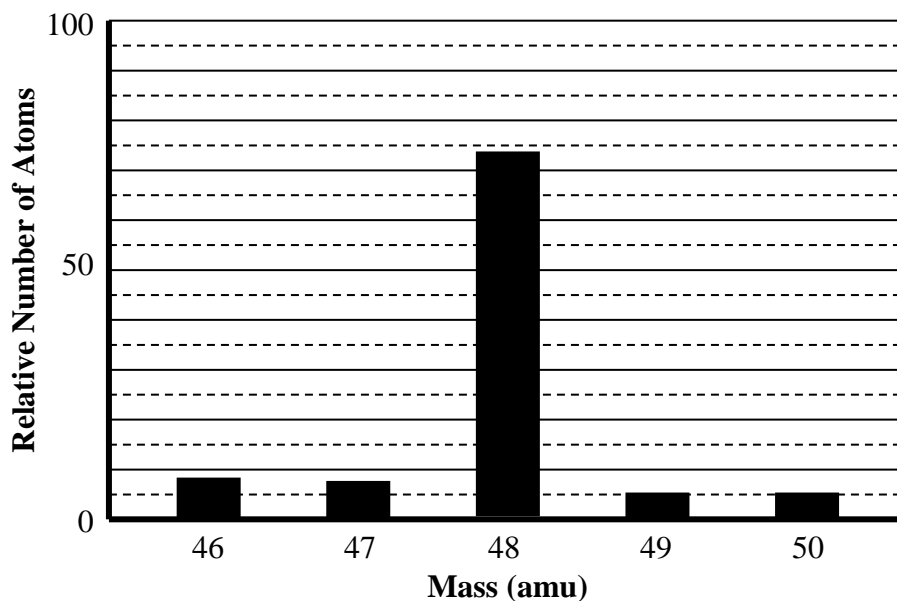
I - Basic Skills:

- 1) Describe how each of the following affects the reaction rate: (a) the physical state of the reactants, (b) the concentration or surface area of the reactants, (c) the pressures of the reactants, (d) the temperature at which the reactions occurs, (e) the presence of a catalyst.
- 2) Determine the order of a reaction, given the initial rate as a function of concentration of reactants.
- 3) For a first order reaction, calculate:
 - (a) the concentration of reactant after a given time, knowing its original concentration and the rate constant.
 - (b) the time needed to drop the concentration by a given amount, knowing the rate constant.
- 4) Given the half-life or the rate constant for a first order reaction, calculate the other.
- 5) Given two of the three quantities: ΔH , E_a , E_a' (activation energy for the reverse reaction), calculate the other quantity.
- 6) Use the Arrhenius equation $[\ln(k_2/k_1) = -E_a/R (1/T_2 - 1/T_1)]$ to calculate any one of the five quantities: k_1 , k_2 , T_1 , T_2 , E_a , knowing or having calculated the other four quantities.
- 7) Determine whether a proposed mechanism for a reaction is consistent with the observed rate expression.
- 8) Be able to differentiate between acid-base catalysis, surface catalysis, and enzyme catalysis.
- 9) Be able to describe how a class of catalyst is usually able to modify a reaction.

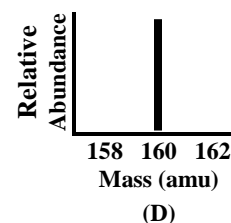
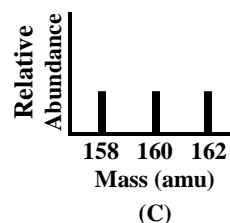
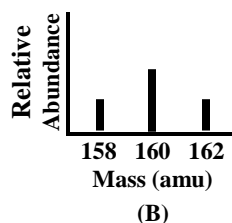
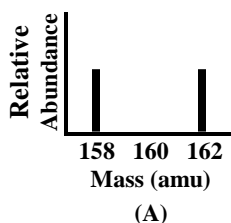
II - Lessons and Assignments Schedule:**THE START OF 2nd SEMESTER MATERIAL!**Day 76 [Mon. 1/11] **Read** 14.1 and 14.2**HW 14.1:** 14.2, 14.20, 14.22, 14.28Lesson-14.1 I **Reaction Rates:** LO 4.2
 II **Rate Laws:**Show me the calculations for $\ln[X]$ & $1/[X]$ in order to see which one linearizes the data.Day 77 [Tue. 1/12] **Read** 14.3 and 14.4**HW 14.2:** 14.4, 14.30a&b, 14.34, 14.36, 14.44Lesson-14.2 III **Change of Concentration with Time:** LO 4.1, LO 4.4, LO 4.2, LO 4.3**The Return and Review of the 1st SEMESTER EXAM**Day 78 [Wed. 1/13] **Read** 14.5 and 14.6**HW 14.3:** 14.46b&c, 14.48, 14.50, 14.54, 14.56Lesson-14.3 IV **Temperature and Rate:** LO 4.1, LO 4.5, LO 4.6Day 79 [Thu. 1/14] **Read** 14.7**HW 14.4:** 14.9, 14.62, 14.64, 14.66, 14.68Lesson-14.4 V **Reaction Mechanism:** LO 4.4, LO 4.5, LO 4.6, LO 4.7Day 80 [Fri. 1/15] **HW 14.5:** 14.10, 14.95, 14.101Lesson-14.5 VI **Catalysis:** LO 4.8, LO 4.9

Student Activity-1: Determining an Element's Average Atomic Mass

- (1) A sample of an element was run through a mass spectrometer resulting in the data printout below. (a) Calculate the average atomic mass for the element sample. (b) Based on your results, predict the identity of the element.



- (2) A sample of pure elemental bromine (Br_2) was run through a mass spectrometer. Atomic bromine is known to have only two isotopes, Br-79 and Br-81 and an atomic mass of 79.904 amu. Which of the following mass spectrum printouts for Br_2 is most consistent with this information? **Remember that elemental bromine is Br_2 !**

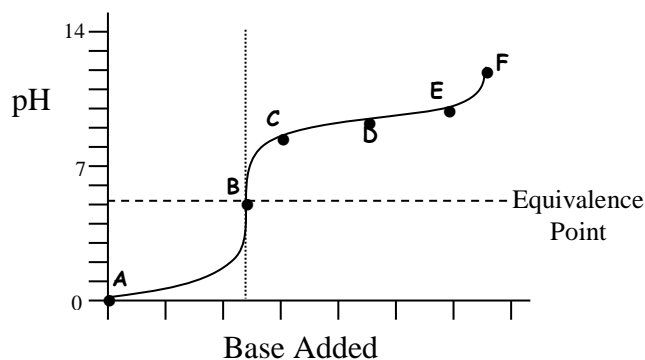


I - Basic Skills:

- 1) Given the balanced equation for a reaction involving gases, write the equilibrium expression for K_c or K_p .
- 2) For a given equation, calculate the value of K_c (K_p), knowing: (a) the equilibrium concentrations (partial pressures) of all species, or (b) the original concentrations (partial pressures) of all species and the equilibrium concentration (partial pressure) of one species.
- 3) Given the value of K_c (K_p), predict: (a) whether the equilibrium lies to the left or to the right. (b) The equilibrium concentration (partial pressure) of one species, given those of all other species. (c) The equilibrium concentrations (partial pressures) of all species, given their original concentrations (partial pressures).
- 4) Given one of the two quantities, K_c or K_p , for a system involving gases at a specific temperature, calculate the other quantity using the equation: $K_p = K_c[RT]^{\Delta n \text{ of gas}}$.
- 5) Using Le Chatelier's Principle, predict the effect of a change in the number of moles, volume, or temperature upon the position of a system at equilibrium.
- 6) Use the van't Hoff equation [$\ln(K_2/K_1) = -\Delta H_R/R (1/T_2 - 1/T_1)$] to calculate any one of the five quantities: K_1 , K_2 , T_1 , T_2 , ΔH_R , knowing or having calculated the other quantities
- 7) For a given system, calculate the value Q_c (Q_p) knowing the concentrations of all species.
- 8) Given or having calculated the value of the Q_c (Q_p) for a system as well as the K_c (K_p) for the same system, determine the direction the system will shift to reach equilibrium.

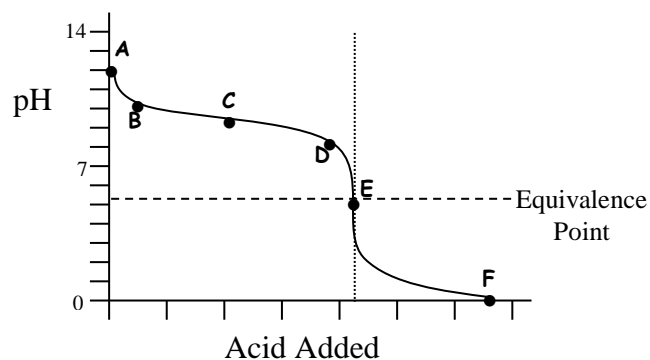
II - Lessons and Assignments Schedule:Day 81 [Tue. 1/19] **Read** 15.1 to 15.5**HW 15.1:** 15.2, 15.3, 15.14ab&e, 15.16, 15.18, 15.22, 15.28, 15.34Lesson-15.1 **I** Systems at Equilibrium: LO 6.1, LO 6.7**II** Calculating the Equilibrium Constant: LO 6.5**Lab-16:** CatalysisDay 82 [Wed. 1/20] **Read** 15.6**HW 15.2:** 15.6, 15.36, 15.38, 15.40, 15.44, 15.48Lesson-15.2 **III** Applications of the Equilibrium Constant: LO 6.2, LO 6.4, LO 6.6, LO 6.10Day 83 [Thu. 1/21] **Read** 15.7**HW 15.3:** 15.8, 15.52, 15.54, 15.56, 15.83-don't answer the question as written instead use the answers in the book for K_p at 850°C and 1200°C and the van't Hoff equation to calculate the ΔH_R for the reactionLesson-15.3 **IV** Le Chatelier's Principle: LO 6.2, LO 6.3, LO 6.8, LO 6.9Day 84 [Fri. 1/22] **TBA TBA TBA TBA TBA TBA**Day 85 [Mon. 1/25] **Lab-20:** The K_{sp} of a Slightly Soluble Salt: Day-1 [$\frac{1}{2}$ period]**MID-YEAR EXAMINATION PERIOD [TUESDAY 1/26 TO FRIDAY 1/29]**You have already taken your 1st semester exam in class and so will not be tested during the school's mid-year examination period.Day 86 [Mon. 2/1] **Review for TEST-5** (Chapters 14-15)Day 87 [Tue. 2/2] **Lab-16:** Chemical KineticsDay 88 [Wed. 2/3] **TEST-5 (day-1) multiple choice:**Day 89 [Thu. 2/4] **TEST-5 (day-2) free response:**

ADDITIONAL EXAMPLES OF TITRATION CURVES FROM PAGE 34



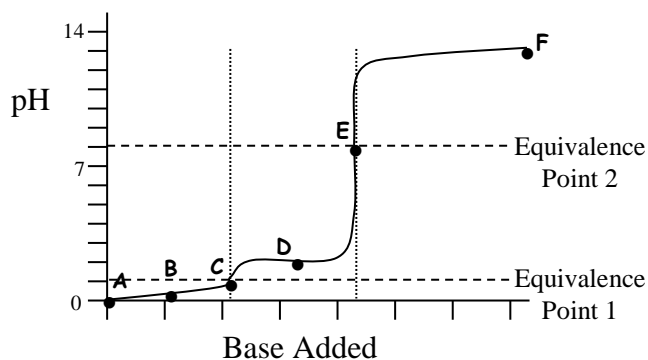
Adding a Weak Base to a Strong Acid
[example: NH_3 and HClO_3]

- A:** H^+ and ClO_3^- dominate
B: NH_4^+ and ClO_3^- dominate
C: NH_4^+ , $\text{NH}_3 \uparrow$, $\text{OH}^- \uparrow$, $\text{ClO}_3^- \downarrow$
D: $\text{NH}_4^+ = \text{NH}_3$ [buffer region], $\text{OH}^- \uparrow$, $\text{ClO}_3^- \downarrow$
E: $\text{NH}_3 \uparrow$, $\text{OH}^- \uparrow$, $\text{ClO}_3^- \downarrow$, $\text{NH}_4^+ \downarrow$,
F: NH_3 and OH^- dominate, $\text{ClO}_3^- \downarrow$, $\text{NH}_4^+ \downarrow$



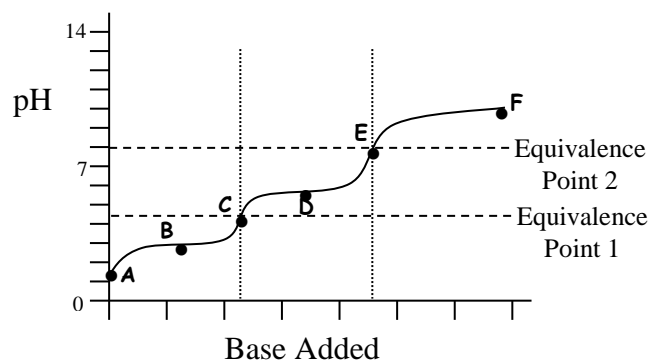
Adding a Strong Acid to a Weak Base
[example: HNO_3 and NaHCO_3]

- A:** Na^+ and HCO_3^- dominate
B: Na^+ dominate, $\text{NO}_3^- \uparrow$, $\text{HCO}_3^- \downarrow$, and $\text{CO}_2 \uparrow$
C: Na^+ and NO_3^- dominate, ...
 ... $\text{HCO}_3^- = \text{CO}_2$ [buffer region]
D: Na^+ and NO_3^- dominate
E: Na^+ and NO_3^- dominate, $\text{H}^+ \uparrow$
F: H^+ and NO_3^- dominate, $\text{Na}^+ \downarrow$



Adding a Strong Base to a Strong Polyprotic Acid
[example: NaOH and H_2SO_4]

- A:** H^+ and HSO_4^- dominate
B: HSO_4^- dominate, $\text{Na}^+ \uparrow$, $\text{H}^+ \downarrow$
C: Na^+ and HSO_4^- dominate
D: Na^+ dominate, $\text{HSO}_4^- \downarrow = \text{SO}_4^{2-} \uparrow$ [buffer region]
E: Na^+ and SO_4^{2-} dominate
F: Na^+ and OH^- dominate, $\text{SO}_4^{2-} \downarrow$



Adding a Strong Base to a Weak Polyprotic Acid
[example: NaOH and $\text{H}_2\text{C}_2\text{H}_2\text{O}_4$]

- A:** $\text{H}_2\text{C}_2\text{H}_2\text{O}_4$ dominate
B: $\text{Na}^+ \uparrow$, $\text{H}_2\text{C}_2\text{H}_2\text{O}_4 \downarrow = \text{HC}_2\text{H}_2\text{O}_4^- \uparrow$ [buffer region]
C: Na^+ and $\text{HC}_2\text{H}_2\text{O}_4^-$ dominate
D: Na^+ dominate, ...
 ... $\text{HC}_2\text{H}_2\text{O}_4^- \downarrow = \text{C}_2\text{H}_2\text{O}_4^{2-} \uparrow$ [buffer region]
E: Na^+ and $\text{C}_2\text{H}_2\text{O}_4^{2-}$ dominate
F: Na^+ and OH^- dominate, $\text{C}_2\text{H}_2\text{O}_4^{2-} \downarrow$

I - Basic Skills:

- 1) Write equations for Bronsted-Lowry reactions and identify conjugate acid-base pairs.
- 2) Predict whether the equilibrium for a Bronsted-Lowry reaction lies more to the left or more to the right based on the relative strengths of the two Bronsted-Lowry bases in the system.
- 3) Calculate the pH of a solution given or having calculated its $[H^+]$ or its $[OH^-]$.
- 4) Calculate the $[H^+]$ or $[OH^-]$ of a solution given or having calculated its pH.
- 5) Calculate the pH or the $[H^+]$ for a strong acid or strong base solution given its concentration.
- 6) Calculate K_a for a weak acid (HA), given the $[H^+]$ or pH of a solution prepared by dissolving the weak acid (HA) in water to a known concentration.
- 7) Given the initial concentration of a weak acid (HA) and the value of K_a , calculate $[H^+]$ or pH.
- 8) Relate the value of K_b for a weak base (B) to K_a of its conjugate acid (HB^{1+}).
- 9) Given the original concentration of a weak base (B) and the value of K_b , calculate $[OH^-]$ or pH.
- 10) Identify the acid and base among the reactants in a Lewis acid-base reaction.
- 11) Write reaction equations for Lewis acid-base reactions.
- 12) Given a hydrated metallic cation (complex ion), write its ionization reaction in water.
- 13) Given the metal cation and the Lewis base, write the formation reaction for the complex ion.
- 14) Compare the strengths of acids (binary, oxy, or carboxylic) to their chemical structures.
- 15) Identify a salt as acidic, basic, or neutral.
- 16) Write the hydrolysis reaction (molecular or net ionic) of an acidic salt or a basic salt in water.

II - Lessons and Assignments Schedule:

Day 90 [Fri. 2/5]		Read 16.1 and 16.3 HW 16.1: 16.2, 16.14, 16.16, 16.20a&b, 16.26, 16.30a&b
Lesson-16.1	I	<u>Arrhenius and Bronsted-Lowry Acids and Bases:</u> LO 3.7, LO 6.15, LO 6.17
	II	<u>Autoionization of Water:</u>
Day 91 [Mon. 2/8]		Read 16.4 and 16.5 HW 16.2: 16.34, 16.40, 16.44, 16.46, 16.48
Lesson-16.2	III	<u>pH and the pH Scale & pOH and the pOH Scale:</u>
	IV	<u>Strong Acids and Bases:</u> LO 2.2
Day 92 [Tue. 2/9]		Lab-17: <u>Chemical Equilibrium and Le Chatelier's Principle</u>
Day 93 [Wed. 2/10]		Read 16.6 and 16.7 HW 16.3: 16.3, 16.4, 16.54, 16.56, 16.58, 16.60, 16.118
Lesson-16.3	V	<u>Weak Acids:</u> LO 2.2, LO 6.5, LO 6.6, LO 6.16, LO 6.19
Day 94 [Thu. 2/11]		Read 16.8 and 16.9 HW 16.4: 16.72, 16.74, 16.76, 16.78, 16.80, 16.82
Lesson-16.4	VI	<u>Comparing Strong and Weak Acid Solutions:</u> LO 6.12
	VII	<u>Weak Bases:</u> LO 2.2, LO 6.6, LO 6.16
	VIII	<u>The Relationship between K_a and K_b:</u>
Day 95 [Fri. 2/12]		Read 16.10 HW 16.5: Homework Worksheet
Lesson-16.5	IX	<u>Lewis Acids and Bases:</u>
	X	<u>Complex Ions:</u>

HAPPY FEBRUARY BREAK!

Day 96 [Mon. 2/22] HW 16.6: 16.94a&c, 16.96, 16.98, 16.100, 16.102

Lesson-16.6 XI Acid-Base Character and Chemical Structure: LO 2.2

Day 97 [Tue. 2/23] HW 16.7: 16.84a, 16.86abc&d, 16.88, 16.112

Lesson-16.7 XII Acid-Base Properties of Salts: LO 6.16

Lab-18: Hydrolysis of Salts

Type of Reaction	K_{eq} for Reaction	Example Reactions	K_{eq} for Example Reactions
Strong acid-strong base	$1/K_w$	$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{HOH}$	1.0×10^{14}
Weak acid-strong base	K_a/K_w	$\text{HF} + \text{KOH} \rightarrow \text{K}^+\text{F}^- + \text{HOH}$	6.8×10^{10}
Strong acid-weak base	K_b/K_w	$\text{HCl} + \text{NH}_3 \rightarrow \text{NH}_4^+ + \text{Cl}^-$	1.8×10^9

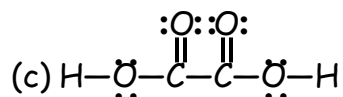
[Answers to student activity from page 20]

(a) It lost mass because the original sample was a hydrate and the water of hydration was released.

(b) HCO_2

(c) $\text{H}_2\text{C}_2\text{O}_4$ - It is oxalic acid

(d) $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$



[Answers to student activity from page 28]

(1) (a) ≈ 47.9 amu

(b) titanium

(2) B

I - Basic Skills:

- 1) Calculate the pH of a solution mixture in which the components contain a common ion.
- 2) Given the composition of a buffer system (HA & A^{1-} or B & HB^{1+}), determine its pH before and after the addition of known amounts of strong acid or strong base.
- 3) Calculate the pH throughout the titration of a ...
(a) strong acid and strong base (b) weak acid and strong base (c) strong acid and weak base
- 4) Determine the dominate species present throughout the titration of a ...
(a) strong acid and strong base (b) weak acid and strong base (c) strong acid and weak base
- 5) Identify the titration curve of a ...
(a) strong acid and strong base (b) weak acid and strong base (c) strong acid and weak base
- 6) Given either the solubility of an ionic compound in water or its K_{sp} , calculate the other.
- 7) Given its K_{sp} , determine the solubility of an ionic compound in a solution with a common ion.
- 8) Given its K_{sp} , determine the solubility of an ionic compound in a solution of given pH.
- 9) Given its K_{sp} , determine the solubility of an ionic compound in which a complex ion is formed.
- 10) Given the formation constant (K_f) for a complex ion, relate the concentrations of free cation (Lewis acid), complex ion and ligand (Lewis base) in solution.
- 11) Use the rules of solubility to determine the presence or absence of an ion in a solution.
- 12) Given, or having calculated, the ion-product (Q_{sp}) for an ionic electrolyte in solution, determine whether or not precipitation will occur.

II - Lessons and Assignments Schedule:Day 98 [Wed. 2/24] **Read** 17.1 and 17.2**HW 17.1:** 17.10, 17.12, 17.14, 17.16**Lesson-17.1** I Acid-Base Solutions and the Common Ion Effect: LO 6.4, LO 6.16Day 99 [Thu. 2/25] **Read** 17.3**HW 17.2:** 17.2, 17.20a, 17.24, 17.26, 17.30**Lesson-17.2** II Buffers: LO 6.16, LO 6.17, LO 6.20Day 100 [Fri. 2/26] **Read** 17.4 and 17.5**HW 17.3:** 17.46b, 17.48a, 17.50, 17.52a**Lesson-17.3** III Solubility Equilibria of Ionic Solids:
A) Solubility of Ionic Solids in Pure Water: LO 6.2, LO 6.5, LO 6.6, LO 6.21, LO 6.22Day 101 [Mon. 2/29] **Read** 17.6**HW 17.4:** 17.7b&c, 17.52b&c, 17.88**Lesson-17.4** III Solubility Equilibria of Ionic Solids:
B) Solubility of Ionic Solids in Salt Solutions: LO 6.2, LO 6.23Day 102 [Tue. 3/1] **Lab-19:** BuffersDay 103 [Wed. 3/2] **HW 17.5:** 17.7a, 17.8, 17.54a&b, 17.56**Lesson-17.5** III Solubility Equilibria of Ionic Solids:
C) Solubility of Ionic Solids and pH: LO 6.2, LO 6.23Day 104 [Thu. 3/3] **HW 17.6:** Homework Worksheet**Lesson-17.6** III Solubility Equilibria of Ionic Solids:
D) Solubility of Ionic Solids that form Complex Ions: LO 6.2Day 105 [Fri. 3/4] **HW 17.7:** 17.62b, 17.66b, 17.70, 17.95**Lesson-17.7** IV Precipitation Reactions: LO 6.4

Day 106 [Mon. 3/7] HW 17.8: Homework Worksheet

Lesson-17.8 V Acid-Base Titrations: Choosing an Appropriate Indicator & Titration Problems: LO 6.11,

Day 107 [Tue. 3/8] HW 17.9: Homework Worksheet

Lesson-17.9 VI Plotting an Acid-Base Titration Curve: (A) a Strong Acid with a Strong Base: LO 6.13

Lab-20: The K_{sp} of a Slightly Soluble Salt: Day-2 [$\frac{1}{2}$ period]

Day 108 [Wed. 3/9] HW 17.10: Homework Worksheet

Lesson-17.10 VI Plotting an Acid-Base Titration Curve: (B) a Weak Acid with a Strong Base: LO 6.13

Day 109 [Thu. 3/10] Student Activity: Plotting an Acid-Base Titration Curve (in Chapter-17 handout)

Lesson-17.11 VI Plotting an Acid-Base Titration Curve: (C) a Weak Base with a Strong Acid: LO 6.13

Day 110 [Fri. 3/11] TBA TBA TBA TBA TBA TBA

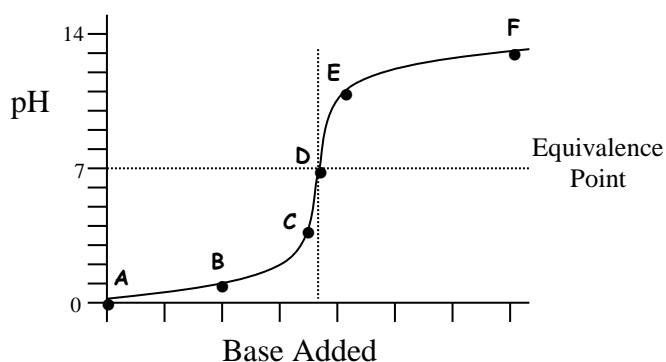
Day 111 [Mon. 3/14] Review for TEST-6 (Chapters 16-17)

Day 112 [Tue. 3/15] Lab-21: Acid-Base Titration [2 periods]

Day 113 [Wed. 3/16] TEST-6 (day-1) multiple choice:

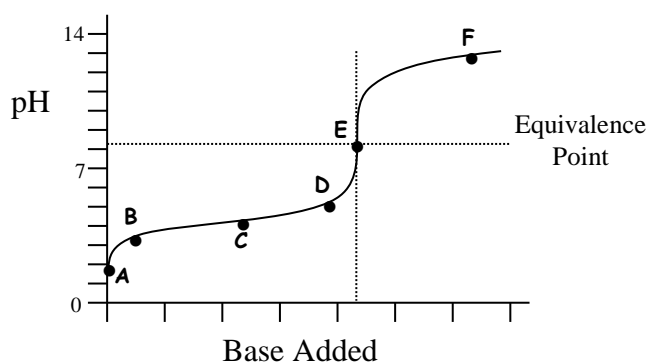
Day 114 [Thu. 3/17] TEST-6 (day-2) free response:

Type of Reaction	pH at Equivalence Point	Appropriate Acid-Base Indicator	Color Change and pH Range of Color Change
Strong Acid and Weak Base	Below 7	methyl orange	red to orange (3.2-4.4)
		bromocresol green	yellow to blue (3.8-5.4)
Strong Acid and Strong Base	7	bromthymol blue	yellow to blue (6.0-7.6)
		litmus	red to blue (5.5-8.2)
Weak Acid and Strong Base	Above 7	phenolphthalein	colorless to pink (8.2-10)
		thymol blue	yellow to blue (8.0-9.6)



Adding a Strong Base to a Strong Acid
[example: HCl and NaOH]

- A: H^+ and Cl^- dominate
- B: Cl^- dominate, $H^+ \downarrow$, $Na^+ \uparrow$
- C: Na^+ and Cl^- , dominate and $H^+ \downarrow$
- D: Na^+ and Cl^- dominate
- E: Na^+ dominates and $Cl^- \downarrow$, $OH^- \uparrow$
- F: Na^+ and OH^- dominate, $Cl^- \downarrow$



Adding a Strong Base to a Weak Acid
[example: HF and KOH]

- A: HF dominates
- B: $HF \downarrow$, $K^+ \uparrow$, $F^- \uparrow$
- C: $HF = F^-$ [buffer region], $K^+ \uparrow$
- D: K^+ and F^- dominate and $HF \downarrow$
- E: K^+ and F^- dominate
- F: K^+ dominate and $OH^- \uparrow$, $F^- \downarrow$

SEE PAGE 30 FOR ADDITIONAL EXAMPLES OF TITRATION CURVES

I - Basic Skills:

- 1) Be able to distinguish between a reversible and an irreversible process.
- 2) Be able to define the terms: (1) entropy and (2) microstate and state the relationship between them both qualitatively and quantitatively.
- 3) Determine the ΔH_{fusion} , T_{fusion} , or ΔS_{fusion} , given the other two.
- 4) Determine the $\Delta H_{\text{vaporization}}$, $T_{\text{vaporization}}$, or $\Delta S_{\text{vaporization}}$, given the other two.
- 5) Be able to apply Trouton's rule to the vaporization of a substance at its normal boiling point.
- 6) Use thermodynamic data (Appendix C in textbook) to calculate ΔH and ΔS° for a reaction.
- 7) Use the Gibbs equation [$\Delta G = \Delta H - T\Delta S$] to calculate the free energy change for a reaction.
- 8) Be able to define the terms:
 - (a) spontaneous process: a process that is able to proceed in a given direction without needing to be driven by any outside source of energy. (*thermodynamically favored*)
 - (b) nonspontaneous process: a process that is unable to proceed in a given direction without an outside source of energy. (*not thermodynamically favored*)
- 9) Describe how the signs of ΔH , ΔS° , and ΔG° relate to the spontaneity of a reaction.
- 10) Given or having calculated ΔH and ΔS° , determine the temperature at which a reaction is at equilibrium at 1 atm.
- 11) State the relative dependency of the values of ΔH° , ΔS° , and ΔG° on temperature, pressure, and concentration changes.
- 12) Given two of the following; ΔG° , T , or K , calculate the third [$\Delta G^\circ = -RT \ln K$].
- 13) Given three of the following; ΔG , ΔG° , T , or Q , calculate the fourth. [$\Delta G = \Delta G^\circ + RT \ln Q$]

II - Lessons and Assignments Schedule:Day 115 [Fri. 3/18] **Read** 19.3**HW 19.1:** 19.10, 19.12, 19.14, 19.20, 19.22Lesson-19.1 I Spontaneous Processes: LO 5.12II Entropy and the Second Law of Thermodynamics: LO 2.15, LO 5.12Day 116 [Mon. 3/21] **Read** 19.4 to 19.6**HW 19.2:** 19.3, 19.28, 19.32, 19.38, 19.40Lesson-19.2 III The Molecular Interpretation of Entropy: LO 5.12Day 117 [Tue. 3/22] **Lab-22:** Redox Titration [2 periods]Day 118 [Wed. 3/23] **Read** 19.7

... 19.70

HW 19.3: 19.42, 19.44ab&c, 19.50, 19.54a&d, 19.64, 19.67a-(show work), ...↑Lesson-19.3 IV Entropy Changes in Chemical Reactions: LO 5.12V Gibb's Free Energy: LO 5.13, LO 5.14Day 119 [Thu. 3/24] **HW 19.4:** 19.6, 19.74, 19.76a&c, 19.102Lesson-19.4 VI Free Energy and Equilibrium: LO 5.15, LO 5.16, LO 5.17, LO 5.18, LO 6.25Student Activity-5: The Thermodynamic Favorability of Reactions (in Chapter-19 handout)

Your second semester exam will take place during the last five days of class.

[Do not be absent! Do not be absent! Do not be absent! Do not be absent! Do not be absent!]

Day-1, Tuesday, June 9th: Chapters 15 & 19

Review Questions from Chapter 15:

15, 21, 27, 29, 45, 47, 49, 51, 59, 71

Review Questions from Chapter 19:

7, 13, 19, 21, 25, 27, 31, 39, 43, 47d, 51, 73

Day-2, Wednesday, June 10th: Chapters 16 & 17

Review Questions from Chapter 16:

25, 31, 39, 43c&d, 47, 53, 61a, 75, 85, 89

Review Questions from Chapter 17:

11a&b, 13a&b, 23, 31, 37a, 39, 49, 51a&b, 55, 59, 61

Day-3, Thursday, June 11th: Chapter 14 & 20

Review Questions from Chapter 14:

8, 27, 43, 55, 94

Review Questions from Chapter 20:

13, 15, 19c&f, 23, 31, 33c&d, 37, 41, 45, 49, 51a, 59, 63, 85, 87, 89, 101

Also be able to predict the products and write a balanced equation for the electrolysis of the following types of aqueous solutions: an acid [ex: HCl, H₂SO₄], a base [ex: KOH], a salt [ex: NaCl, Cu(NO₃)₂, K₂SO₄, SnBr₂]

Day-4, Friday, June 12th: Post AP Exam Topics and Laboratory Activities

Review Questions from Chapter 11:

Review Questions from Chapter 13:

Review Questions from Chapter 25:

Review Questions from Laboratory Activities:

Day-5, Monday, June 15th: The Return and Review of the Second Semester Exam

I - Basic Skills:

- 1) Given the formula of a species, determine the oxidation number (state) of each atom.
- 2) Given a redox equation, select the oxidizing agent and the reducing agent.
- 3) Given the formulas of the reactants and the product, balance a redox equation by the half-equation method.
- 4) Given a balanced equation or half-equation in an acidic solution, write the corresponding equation in a basic solution.
- 5) Contrast electrolytic cells with voltaic (galvanic) cells; identify the anode and the cathode in either type of cell and indicate the direction of the current flow (or electron flow) through all parts of the cell.
- 6) Use standard electrode potentials (Table 20.1 and Appendix E in textbook) to:
 - (a) compare the relative strengths of different oxidizing or reducing agents.
 - (b) calculate a cell voltage at standard concentrations.
 - (c) decide whether or not a given redox reaction will occur spontaneously under standard concentrations.
- 7) Use the Nernst equation [$E_{\text{cell}} = E_{\text{cell}}^{\circ} - (RT/nF)\ln Q$] to calculate:
 - (a) the voltage of a cell, given E_{cell}° and the concentrations of all species.
 - (b) the concentration of one species, given the concentrations of all other species involved as well as E_{cell} and E_{cell}° .
- 8) Predict the products when an ionic solute is electrolyzed in an aqueous solution.
- 9) Relate the number of electrons or coulombs passing through an electrolytic cell to the amounts of products formed at the electrodes.
- 10) Relate the amount of energy used in electrolysis (joules or kilowatt hours) to the number of coulombs and volts.

II - Lessons and Assignments Schedule:

Day 120 [Mon. 3/28] **Read** 20.1 and 20.2

HW 20.1: Homework Worksheet

Lesson-20.1 **I** **Redox Reactions and Oxidation States:**

Day 121 [Tue. 3/29] **Read** 20.3

HW 20.2: Homework Worksheet

Lesson-20.2 **II** **Balancing Redox Reactions:**

Lab-23: **Reactions of Acids and Bases** [1 period]

Day 122 [Wed. 3/30] **HW 20.3:** 20.16, 20.18bc&f, 20.20abc&d

Lesson-20.3 **II** **Balancing Redox Reactions (Continued):**

Day 123 [Thu. 3/31] **Read** 20.4

HW 20.4: 20.18de&g, 20.20e&f, 20.22, 20.24

Lesson-20.4 **III** **Voltaic (Galvanic) Cells:** LO 3.12, LO 3.13, LO 5.15

Day 124 [Fri. 4/1] **Read** 20.5

HW 20.5: 20.26, 20.28, 20.32, 20.34, 20.36

Lesson-20.5 **IV** **Cell EMF Under Standard Conditions:** LO 3.12, LO 3.13

Day 125 [Mon. 4/4]	Read 20.6
	HW 20.6: 20.48, 20.52, 20.54, 20.64ab&d
Lesson-20.6	V <u>Free Energy and Redox Reactions:</u>
	VI <u>Concentration Cells:</u>
Day 126 [Tue. 4/5]	Lab-24: <u>Spectrophotometry</u> [2 periods]
Day 127 [Wed. 4/6]	Read 20.9
	HW 20.7: 20.56, 20.64c, 20.58, 20.60, 20.99
Lesson-20.7	VII <u>Cell EMF under Nonstandard Conditions:</u> LO 3.12, LO 3.13
	<u>Student Activity:</u> Equilibrium Changes in a Voltaic Cell (in Chapter-20 handout)
Day 128 [Thu. 4/7]	HW 20.8: Homework Worksheet
Lesson-20.8	VIII <u>Electrolysis and the Electrolytic Cell:</u> LO 5.15
Day 129 [Fri. 4/8]	HW 20.9: 20.84, 20.86, 20.90, 20.106-for part b, list two reasons
Lesson-20.9	XI <u>Quantitative Aspects of Electrolysis:</u> LO 3.12, LO 3.13, LO 5.15

THE END OF MATERIAL COVERED ON THE AP EXAM!

Day 130 [Mon. 4/11]	Review for AP test
Day 131 [Tue. 4/12]	Review for AP test
	Review for TEST-7 (Chapters 19&20)
Day 132 [Wed. 4/13]	<u>TEST-7 (day-1) multiple choice:</u>
Day 133 [Thu. 4/14]	<u>TEST-7 (day-2) free response:</u>
Day 134 [Fri. 4/15]	Review for AP test

[Suggested topics for the student assignment from page 44 (Curricular Requirement 4)]

(1) Acid Rain in the Adirondacks, (2) Is Climate Change Affecting the Intensity of Storms? (e.g. Sandy or Irene in NYS), (3) The Science and Politics of Hydrofracking in NYS, (4) Is Ethanol an Effective Fuel Additive?, (5) Drug Doping in Sports, (6) Are Home Improvement Store Plants Killing Bees?, (7) The Controversy of Water Fluoridation, (8) Is There a Link between Childhood Vaccinations and Autism?, (9) The Science and Politics of HPV Vaccination, (10) Does Coffee Cure or Kill?, (11) The Efficacy of Green Household Cleaners, (12) Genetically Modified Foods: The Pros and Cons, (13) The Controversy of High Fructose Corn Syrup, (14) PCB's and Dredging the Hudson River, (15) Did They Really Have to Ban Mercury Thermometers from Public Schools?

Pre-AP Week:Day 135 [Mon. 4/18] **Review** for AP testDay 136 [Tue. 4/19] **Review** for AP testDay 137 [Wed. 4/20] **Review** for AP testDay 138 [Thu. 4/21] **Review** for AP testDay 139 [Fri. 4/22] **Review** for AP test**HAPPY APRIL BREAK!**

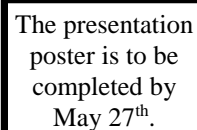
You probably will want to study for the AP examination.

AP EXAMINATION PERIOD**[MONDAY 5/2 TO FRIDAY 5/13]**

During the two weeks of the AP exams, many of you will be in and out of class. During this time, the normal class schedule will be suspended. (1) You and your lab partner will select and research a topic that connects chemistry to a societal issue (see page 38 for ideas) and prepare a joint research poster (see page 49 for directions). (2) You will work on Lab 25: The Electrolytic Cell. (3) You will work on homework assignments HW EP.1-due 5/10 and HW EP.2-due 5/11 (You will need to read section 18.4 in order to complete HW EP.2)

AP Week-1:

	AM	PM
Day 140 [Mon.5/2]	<u>Show Time!</u> <u>CHEMISTRY</u> / Environ. Sci.	Psychology
Day 141 [Tue.5/3]	Spanish Language	Physics1/Art History
Day 142 [Wed. 5/4]	English Lit. & Comp.	Physics2
Day 143 [Thu. 5/5]	Calculus AB / Calculus BC	No Exams at BCHS
Day 144 [Fri.5/6]	US History / Studio Art	European History

Research Topic for Presentation Poster**Lab-25:** The Electrolytic Cell [2 periods]**HW EP.1:** Homework Worksheet (due 5/10)**HW EP.2:** (Read section 18.4) 18.28, 18.30, 18.60, 18.70, 18.72 (due 5/11)


The presentation poster is to be completed by May 27th.

AP Week-2:

	AM	PM
Day 145 [Mon. 5/9]	Biology / Music Theory	Physics C
Day 146 [Tue.5/10]	No Exams at BCHS	French Language
Day 147 [Wed.5/11]	No Exams at BCHS	Macroeconomics
Day 148 [Thu.5/12]	World History	Statistics
Day 149 [Fri. 5/13]	Microeconomics	No Exams at BCHS

Science is a systematic, self-correcting, ever-evolving way of thinking about the world around us. The triumphs and the blunders of one generation of scientists become the foundation upon which the next generation of scientists builds their understanding of the universe. In 1953, the eminent scientist, mathematician, and philosopher Jacob Bronowski wrote in a series of three essays - first given as lectures - and latter to become his famous book, Science and Human Values,

"Most often today's theories flatly contradict those of 1760; many contradict those of 1900. In cosmology, in quantum mechanics, in genetics, in the social sciences, who now holds the beliefs that seemed firm sixty years ago (\approx 1900)? Yet the society of scientists has survived these changes without a revolution, and honors the men whose beliefs it no longer shares. No one has recanted abjectly at a trial before his colleagues. The whole structure of science has been changed, and no one has been either disgraced or deposed."

"But human search and research is a learning by steps of which none is final, and the mistakes of one generation are rungs in the ladder, no less than their correction by the next."

"Science at last respects the scientist more than his theories; for by its nature, it must prize the search above the discovery, and the thinking (and with it the thinker) above the thought." [Notice the non-inclusive gender language of the 1950's.]

Be aware that, even though science is presented in class as a set of neat and concise ideas, the actual activity of science is anything but neat and concise.

"Science is not a mechanism but a human progress, and not a set of findings but a search for them." (Jacob Bronowski, Science and Human Values)

It is important to recognize that popular culture often misunderstands science. It usually views science as simply "a set of findings" or that the scientific term "theory" (as in ... "a long tested model or explanation that satisfactorily accounts for a certain set of phenomena") and the popular term "theory" (as in ... "I have a thought or idea about something") are equivalent.

Thus, it is necessary that we all attempt to improve our level of scientific literacy as well as our ability to understand how popular culture is influenced by science and how science is influenced by popular culture. Keep in mind that, like the arts and literature, science is a cultural pursuit.

- ☞ This year, you will have an opportunity to read and discuss articles from journals such as Scientific American, the Journal of Chemical Education, and Science or books such as The New Alchemists by Robert Hazen, The Elements by Primo Levi, or The 13th Element by John Emsley.
- ☞ This year, you will have an opportunity to read and discuss poems such as "Paracelsus" by Robert Browning, novels such as Frankenstein by Mary Shelly or Dr. Jekyll and Mr. Hyde by R. L. Stevenson, or plays such as "Inherit the Wind" by Jerome Lawrence and Robert E. Lee.
- ☞ This year, you will have an opportunity to view and discuss images of art works such as "An Alchemist in His Workshop" by David (The Younger) Teniers, "Chemical Still Life" by Taylor Anne Smith or "Carbon Rapture" a series of sculptures made by the University of Keele (UK).
- ☞ This year, you will have an opportunity to listen to and discuss recordings of music such as "Carbon Fourteen" by Sting, "The Elements" by Tom Lehrer, or "the Planets" by Gustav Holst.
- ☞ This year, you will have an opportunity to view and discuss films such as "Race for the Double Helix" (based on the book by James Watson), "Copenhagen" (based on the play by Michael Framing), "October Sky" (based on the book, Rocket Boys, by Homer H. Hickam Jr.), or "Madame Curie" (based on the biography by her daughter Eve Curie). [***At my first NY teaching post, the chemistry laboratory was first designed and then opened by Madame Curie! ©***]

I - Basic Skills:**From Chapter-11:**

- 1) Determine the type of cubic unit cell for a pure metallic solid (simple cubic, face-centered cubic, or body-centered cubic), the cell dimensions, or the atomic radius, given the other two.

From Chapter-13:

- 2) Given appropriate data, calculate the mass percent, ppm, ppb, mole fraction, molarity or molality of a solution.
- 3) Convert between the molarity, molality, mole fraction, and mass percent of a solution.
- 4) Use the equations, $\Delta T_b = k_b m_i$ and $\Delta T_f = k_f m_i$, to obtain
 - (a) the boiling point or freezing point of a nonelectrolyte solution, knowing or having calculated the molality.
 - (b) the molar mass of a nonvolatile nonelectrolyte.
- 5) Compare the colligative properties of electrolytes to those of nonelectrolytes and be able
 - (a) to determine the ideal van't Hoff factor (i) for a solute.
 - (b) to calculate the approximate boiling point or freezing point of an electrolyte solution, knowing or having calculated the molality.

From Chapter-25:

- 6) Given the structural or the molecular formula for an alcohol or a carboxylic acid be able to name it.
- 7) Given the name for an alcohol or a carboxylic acid be able to give its structural or molecular formula.
- 8) Given the names or formulas for the reactants, predict the products for the esterification of an organic acid (carboxylic acid) and an alcohol.

II - Lessons and Assignments Schedule:

Day 150 [Mon. 5/16] **HW PEP.1:** Homework Worksheet

Lesson-1 I The Crystalline Structure of Metallic Solids:

Day 151 [Tue. 5/17] **Read** 13.4

HW PEP.2: 13.5, 13.7, 13.34b, 13.42, 13.48a, 13.50

Lesson-2 II Concentrations of Solutions: LO 2.9

III Conversions between Concentration Units:

Day 152 [Wed. 5/18] **Read** 13.5 and 13.6

HW PEP.3: 13.8, 13.56, 13.58, 13.64, 13.68a, 13.70, 13.72

Lesson-3 IV Colligative Properties:

Day 153 [Thu. 5/19] **HW PEP.4:** Homework Worksheet

Lesson-4 V Alcohols: LO 2.1

Day 154 [Fri. 5/20] **Lab-26:** Unit Cells: The Crystalline Structure of Metallic Solids

Day 155 [Mon. 5/23] **HW PEP.5:** Homework Worksheet

Lesson-5 VI Carboxylic Acids: LO 2.1

Day 156 [Tue. 5/24] **Lab-27:** Freezing Point Depression and Molecular Mass

Day 157 [Wed. 5/25] **HW PEP.6:** Homework Worksheet

Lesson-6 **VII** **Esters:** LO 2.1

Day 158 [Thu. 5/26] **HW PEP.7:** Homework Worksheet

Lesson-7 **VIII** **Esterification:** LO 2.1

Day 159 [Fri. 5/27] **Presentation Poster Due**

Day 160 [Tue. 5/30] **Lab-28:** **Esterification**

2nd Semester Exam Review [See page 36 for suggested review]

Day 161 [Wed. 6/1] **Review: Additional Topics**

Day 162 [Thu. 6/2] **Quiz: Additional Topics**

Day 163 [Fri. 6/3] **2nd Semester Exam Review [See page 36 for suggested review]**

Day 164 [Mon. 6/6] **2nd Semester Exam Review [See page 36 for suggested review]**

Day 165 [Tue. 6/7] **2nd SEMESTER EXAM (Part I:)**

2nd Semester Exam Review [See page 36 for suggested review]

Day 166 [Wed. 6/8] **2nd SEMESTER EXAM (Part II:)**

Day 167 [Thu. 6/9] **2nd SEMESTER EXAM (Part III:)**

Day 168 [Fri. 6/10] **2nd SEMESTER EXAM (Part IV:)**

Day 169 [Mon. 6/13] **The Return and Review of the 2nd SEMESTER EXAM**

- ☞ There are 202 instructional periods scheduled for this year (169 class periods and 33 lab periods). A minimum of 50.5 instructional periods (25% of 202) are dedicated to hands-on laboratory work. [CR5a]
- ☞ You will be expected to keep a lab notebook which will include all of your raw data and observations. In some cases, you will be expected to develop your own experimental lab procedure and/or data tables [all lab procedures must first be approved by your teacher before you begin work]. [CR7]
- ☞ Unlike first-year chemistry, you will be expected to “get results” for some lab activities. Extra time may be required to obtain adequate results. You may have to continue and/or redo some trials for certain lab activities during 9th period (extra help time). All raw data (even rejected data) will be kept in your notebook. [CR7]
- ☞ Any laboratory activity that is designated with a ** symbol means that it must be done in its entirety during 9th period (extra help time). You will be given a window of time to complete the lab. All labs may require extra time during 9th period.
- ☞ Even when you work in pairs or groups during a laboratory activity, each of you will be required to write your own individual lab report. [CR7]
- ☞ If you are absent during lab, make-up work must be done during 9th period.

These are the activities that you will perform before the AP Exam period

****Lab-1: Graphing in Microsoft Excel (2 periods) SP: 2,5**

Purpose:

In this activity, you will learn how to use Microsoft Excel to appropriately graph data. You will analyze sets of graphed data in order to make conclusions.

Lab-2: Accuracy and Precision (1 period) SP: 2,4,5

Purpose:

In this activity, you will determine the accuracy of a graduate cylinder

Lab-3: Density of an Unknown Solid [*A GUIDED INQUIRY LAB*] (2 periods) SP: 2,4,5

Purpose:

In this activity, you will determine the density of an unknown solid. *Since you have completed density determinations in previous science courses, you will generate your own procedure and data tables. Based on your results and other observations, you will attempt to identify the unknown solid.*

Lab-4: Empirical Formula of a Compound (1 period) SP: 1,2,5,6

Purpose:

In this activity, you will experimentally determine the empirical formula for a compound of magnesium and oxygen.

Lab-5: Paper Chromatography (1 period) SP: 1,2,5,6

Purpose:

In this activity, you will separate a mixture using the technique of paper chromatography.

Lab-6: Quantitative Analysis of a Chemical Reaction (2 periods) **SP: 1,2,3,5****Purpose:**

In this activity, you will determine the mole ratio of the solid reactant (iron) and the solid product in a reaction between iron metal and copper (II) chloride. You will write a balanced equation for the reaction.

Lab-7: Qualitative Analysis of a Solution [A GUIDED INQUIRY LAB] (2 periods) **SP: 3,4,5****Purpose:**

In this activity, you will determine the identity of two sets of unknown salt solutions. *After completing Part-I of this lab (set-1), you will generate your own procedure and data table for Part-II (set-2).*

Lab-8: Heat of Neutralization (1 periods) **SP: 2,5****Purpose:**

In this activity, you will determine the molar heat of neutralization of 1.00 M HCl and 1.00 M NaOH in kilojoules.

Lab-9: Law of Dulong and Petit (2 periods) **SP: 2,5,6****Purpose:**

In this activity, you will determine the specific heats of an unknown metal and using your experimental value for specific heat and the Law of Dulong and Petit, you will approximate the molar mass for the unknown metal. *Based on your results and observations, you will attempt to identify the unknown metal. You will evaluate the merits and shortcomings of this method to determine molar mass.*

Lab-10: Precipitation Titration (2 periods) **SP: 2,5****Purpose:**

In this activity, you will determine the molar concentration of chloride ions in a water sample [*Mohr's Method*] by titrating it with a silver nitrate solution of known concentration (**0.10 M**).

Lab-11: Activity Series of Selected Metals (1 period) **SP: 5****Purpose:**

In this activity, you will rank, from highest to lowest, the activity of the eight pure metals based on your experimental observations.

Lab-12: Molar Mass of an Unknown Volatile Liquid (1 periods) **SP: 2,5,6****Purpose:**

In this activity, you will determine the molar mass of an unknown volatile liquid. You will evaluate the merits and shortcomings of this method to determine molar mass.

Lab-13: Activity Series of the Halogens (1 period) **SP: 5****Purpose:**

In this activity, you will rank, from highest to lowest, the activity of the pure halogens - chlorine, bromine, and iodine -based on your experimental observations.

Lab-14: Heat of Solution [A GUIDED INQUIRY LAB] (2 periods) **SP: 2,4,5****Purpose:**

In this activity, you will determine the molar heat of solution for NH_4Cl and for CaCl_2 . *Since this is the third calorimetry lab, you will generate your own procedure and data tables.*

Lab-15: Catalysis (1 period) **SP: 5,6,7****Purpose:**

In this activity, you will determine the effect of catalysts on the reaction rates of chemical reactions. You will evaluate the effectiveness of a homogeneous catalyst versus a heterogeneous catalyst on the reaction.

Lab-16: Chemical Kinetics (2 periods) **SP: 1,2,5,6,7****Purpose:**

In this activity, you will determine the reactant orders and a specific rate constant for a chemical reaction.

Lab-17: Chemical Equilibrium and Le Chatelier's Principle [A GUIDED INQUIRY LAB] (2 periods)**Purpose:****SP: 3,4,5**

In this activity, you will study how an equilibrium system responds to various stresses. *Given an equilibrium system, you will decide on which stresses to test for and then generate your own procedure and data tables.*

Lab-18: Hydrolysis of Salts [A GUIDED INQUIRY LAB] (1 period) **SP: 4,5****Purpose:**

In this activity, you will determine the pH of a number of salt solutions. You will explain the formation of neutral and non-neutral salt solutions by writing their dissolving equations as well as any chemical equations for salt hydrolysis. *You will generate your own procedure and data tables.*

Lab-19: Buffers (2 periods) **SP: 5****Purpose:**

In this activity, you will measure the pH changes that occur when acids and bases are added to buffered and unbuffered solutions.

Lab-20: The K_{sp} of a Slightly Soluble Salt (1 periods) **SP: 2,5****Purpose:**

In this activity, you will determine the K_{sp} of the slightly soluble salt lithium carbonate.

Lab-21: Acid-Base Titration (2 periods) **SP: 2,5****Purpose:**

In this activity, you will determine the molar concentration of an acetic acid [$\text{HC}_2\text{H}_3\text{O}_2$] solution by titrating it with a base [NaOH] of known concentration (**0.15 M**). You will generate a pH curve for the titration of acetic acid and sodium hydroxide in order to estimate the K_a value of acetic acid.

Lab-22: Redox Titration [A GUIDED INQUIRY LAB] (2 periods) **SP: 1,2,3,4,5****Purpose:**

In this activity, you will determine the molar concentration of an unknown solution of KMnO_4 by titrating it with a $\text{Na}_2\text{C}_2\text{O}_4$ solution of known concentration (**0.00400 M**) that has been acidified with sulfuric acid. *This being the second titration lab, you will generate your own procedure and data tables.*

Lab-23: Reactions of Acids and Bases (1 period) SP: 5,6**Purpose:**

In this activity, you will observe the reactions of a strong acid with several metals, a carbonate, and a hydrogen carbonate. You will observe the reaction of a strong base with an ammonium salt. You will observe the reactions of carbon dioxide and calcium oxide with water as well as their effects on the pH of the water. You will observe the reaction of HCl gas with NH₃ gas.

Lab-24: Spectrophotometry (2 periods) SP: 2,5**Purpose:**

In this activity, (1) based on KMnO₄'s absorption spectrum, you will select an appropriate wavelength to plot a standard absorption curve and then (2) you will plot a standard absorption curve at that specific wavelength for a known solution of KMnO₄ and then (3) you will use the curve to determine the molar concentration of an unknown solution of KMnO₄ (*the same unknown solution from lab-22*).

This is the activity that you will perform during the AP Exam period.

Lab-25: The Electrolytic Cell (2 periods) SP: 2,5**Purpose:**

In this activity, you construct an electrolytic cell and you will undertake the electrolysis of an aqueous salt solution. You will construct an electroplating apparatus and use it to electroplate a penny with silver.

These are the activities that you will perform after the AP Exam period.

Lab-26: Unit Cells: The Crystalline Structure of Metallic Solids (1 period) SP: 1,2,5,7**Purpose:**

In this activity, you will use models to construct three cubic unit cells and to verify the relationship between the atomic radius of metals and the edge length of their unit cells.

Lab-27: Freezing Point Depression and Molar Mass Determination (2 periods) SP: 2,5,6**Purpose:**

In this activity, you will determine the molar mass of an unknown sugar. You will evaluate the merits and shortcomings of this method to determine molar mass.

Lab-28: Esterification (1 period) SP: 5**Purpose:**

In this activity, you will prepare an ester and study some of its properties.

In science, as in most endeavors, the facility for making accurate observations and recording them in comprehensible fashion is of central importance. The bridge between laboratory observations and subsequent understanding of them is the observer's notebook.

Individual scientists, after some experience with this process, develop methods of record keeping which are uniquely suited to their purposes and inclinations. A criterion one should apply to one's notebook keeping is that another investigator with a reasonable degree of familiarity with the field should be able to reproduce your work using your notebook as a guide.

As a starting point for your record keeping you may find the following observations and suggestions on notebooks useful.

- (1) Good notebook writers acquire their ability through thoughtful practice not inheritance.
- (2) Most useful notebooks have the following characteristics:
 - a) A full name and address of the owner inside or on the front cover.
 - b) An index of experiments near the front.
 - c) Numbered and initialed pages.
 - d) A title for each experiment with a date and a full signature of the investigator.
 - e) A consistent and clear organization or format for the recording of data and observations.
 - f) Concise and legible handwriting.
- (3) Promptness in recording experimental observations in the notebook enhances their accuracy. Ideally the notebook is a record of what you observed and not a record of what you thought you should have observed or of what you observed colored by your expectations. This difficult aspect of recording observations calls for the utmost in personal honesty. If you succeed in being honest, you will find your reasoning and understanding of what you are doing will be continually challenged as you seek the whys and wherefores underlying the observations.
- (4) The notebook should be a bound hard-covered book. Spiral or loose-leaf notebooks are not recommended, and the keeping of experimental observations on loose sheets or paper scraps is not acceptable. If corrections are necessary, these should be made with additional notes rather than erasures.
- (5) All data, calculations, notes and doodles pertinent to the experiment should be in the notebook. A convenient and highly recommended way of accomplishing this end while preserving order is to use the left hand of the pages of the notebook for calculations, tables of physical constants that pertain to the experiment and that are imported from sources such as the CRC, rough copies of chromatograms and/or any other instrument tracings, flowcharts or other schematic drawings, safety cautions, and other reminders, and to use the right hand pages for the actual record of the experiment.
- (6) The right hand side of your notebook should be concise, legible, and have clear and labeled sections. You should skip a line or two in between the sections.
 - a) A title for the experiment with the date and your full signature.
 - b) A brief statement of the objective of the experiment in a section clearly titled "Purpose".
 - c) During the experiment, you should record what you did and what you observed at the time in a section clearly titled "Procedures and Data" (you may want to keep rough notes and data on the left side and then to write a coherent record of what you did soon after, including data and observations, on the right side). Do not rewrite the procedure from the lab handout. Instead tell, in the briefest possible way, what materials you handled and what data and/or observations you obtained, so that someone else familiar with the subject could reproduce your work.

A laboratory notebook should never be confused with a formal laboratory report which is a much more polished piece of work and which contains an introductory section that includes background information about the experiment as well as a final section on the discussion and conclusion of the results.

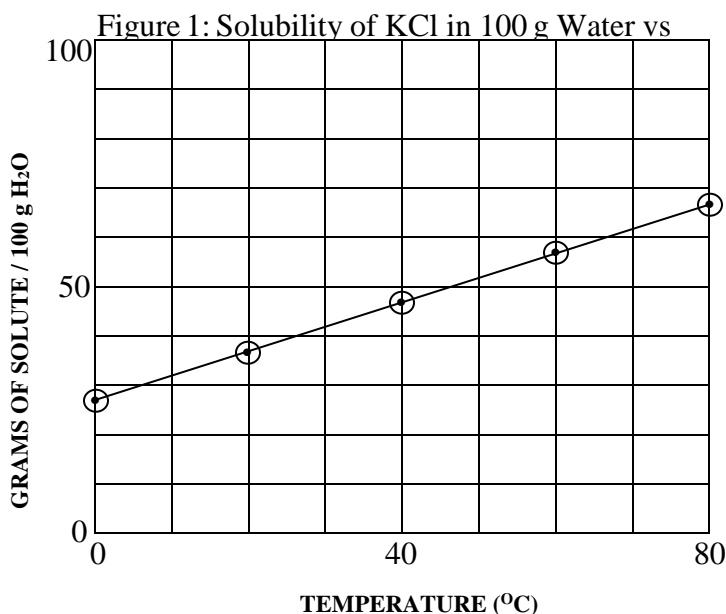
Although there are many different and specific ways to write a formal lab report, there are certain general guidelines that are common to all. For this class, use the following guidelines.

- ☞ Make your report concise. Type your report. On the first page, use the class heading and title your report.
- ☞ Arrange your report logically. Use the following format and subtitle each section with these headings:
 - Introduction: You should have a paragraph which introduces and gives background to the experiment. The last sentence of the introduction section should have a clear statement of the purpose or purposes of the experiment.
 - Materials and Methods: You may list the materials (equipment and supplies) used. You should write the procedure in paragraph form. Be concise, but include enough detail so that anyone could repeat the experiment as you performed it. Use the passive voice in the section of the report.
 - Results and Discussion: In the “Results and Discussion” section, you are to talk about your data and observations. For each and every Table and Figure you include, you must refer to it somehow and someplace in this section. In other words, “walk” the reader through your results. What do your Tables “tell” you that you want the reader to be aware of? Are there any observations or any other information that you want the reader to know about that would help explain the data in the Tables? Did you use any equations to produce results from data? What were the equations? What do your graphs “tell” you that you want the reader to be aware of? Does the plotted data suggest a trend or pattern, what is that trend or pattern, and what might the trend or pattern mean?

Data and observations should be placed in numbered and titled tables.

Table 1: Solubility of KCl in 100 g Water at Various Temperatures	
Temperature (°C)	Mass of KCl/100 g H ₂ O (g)
0	27
20	37
40	47
60	57
80	67

- Graphs should be inserted in this section. Graphs are numbered and titled. All graphs and any other diagrams or pictures are referred to as figures.



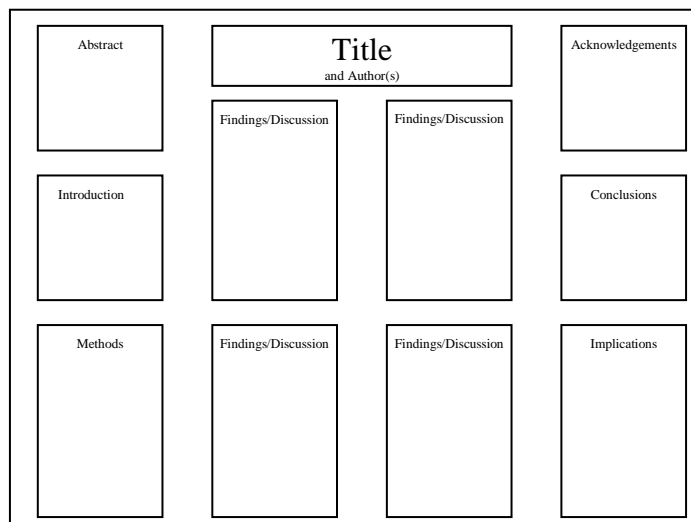
- Conclusion: In this section, you wrap up the report. Did you or did you not accomplish your goal or goals. Did you find something that was unusual or unexpected? In your conclusion, you may pose a question or problem for a future experiment. For example: *“The data suggest that the solubility of KCl in water is directly related to temperature between 0°C and 80°C. (1) Would this relationship continue up to 100°C? (2) It would be interesting to determine whether the solubility of other salts in water also have a direct linear relationship with respect to temperature.”*

How to Create a Poster for Scientific or Literature Research

A poster is a visual communication tool for research but it is not a research document. Unlike a research paper or journal article, a poster is meant to communicate a message quickly. The entire poster should be able to be viewed and understood within about one minute (or less). In order to do this, your poster needs to have a narrow focus and a direct message. For this class, use the following guidelines.

- ☞ Make your poster concise. Use as little text as possible. Instead, use tables and figures (graphs, diagrams, and pictures) when appropriate. Leave empty space – do not fill up every square inch of your poster. Have one main message. A good poster should make people want to stop and look – and hopefully read.
- ☞ Posters can vary in size depending on the venue. For this class you will use a 48in. by 36in. poster.
- ☞ Arrange your poster logically. Use the following general format:
 - Title and Author(s): The title should be large and should be centered at the top of your poster. Your name(s) should be smaller but prominent.
 - Abstract: A short description of the research should be written in paragraph form.
 - Introduction: A brief review of the literature and a clear statement of the purpose – is given (in bullet form if possible).
 - Method and Materials (*for scientific research only*): A brief description of the experimental design – is given (in bullet form if possible).
 - Results/Findings and Discussion: Use tables and figures (graphs, diagrams, or pictures) as much as possible. A brief caption should discuss/explain each graphic used.
 - Conclusion: What did you learn? (*For scientific research*) Did you or did you not accomplish your goal(s)? Was your hypothesis valid? What future research may be necessary?
 - Implications/Applications: This is where you voice your own ideas or opinions. Suggest what potential impact your findings may have on society.
 - Acknowledgements: Identify your sources of information and/or individuals or organizations that significantly helped you.
 - Poster Layout: Organize your poster in columns so it is easy to read. Title and number the different sections so that the viewer knows how to navigate the poster. Make the tables figures large enough so that, from a distance, your poster looks interesting. Remember, a good poster should attract viewers. Include the bare minimum of information that will still communicate your research.

The example on the right shows only one possible design set up. Try different set ups before you glue anything down.



Answers:**Page 56:**

- 1) a) i and v b) vi c) iv d) ii and iii
- 2) a) mixture, homogeneous
b) mixture, homogeneous
c) pure substance
d) mixture, heterogeneous
- 3) Solids have a definite shape and a definite volume. Liquids have an indefinite shape but a definite volume. Solids have an indefinite shape and an indefinite volume.
- 4) a) chemical change
b) physical change
c) physical change
d) chemical change
e) chemical change

Page 66:

- 1) a) kilogram b) Kelvin
- 2) a) centi b) none c) nano
- 3) a) 0.0320 g b) 570 mL c) 90.2°C d) 1.60×10³ m
- 4) 33.7 g
- 5) a) 77°F b) -4°F
- 6) 373 K
- 7) 7) -273°C
- 8) 8) 6.71×10⁸ mi/hr

Page 74:

- 1) 2 2) 1 3) 1 4) 2 5) 4 6) 1
- 7) 2 8) 3 9) 4 10) 3 11) 2

A) Forms of Matter:

Matter is anything that has mass and volume. All matter exists in one of three forms: elements, compounds, or mixtures. Elements and compounds are pure forms of matter which are called substances. Mixtures are impure forms of matter.

1) Elements:

An **element** is a form of matter that cannot be decomposed by either a physical or a chemical reaction. An element is a pure homogeneous substance and is composed of atoms. There are currently over 110 known elements. An element's symbol consists of one or two letters: the first of which is always upper case and if a second is used it is always lower case.

2) Compounds:

A compound is **composed of two or more different elements that are chemically combined (bonded) in a definite ratio**. The formula for a compound consists of the element symbols followed by a subscript which indicates the actual number of atoms that are present in the smallest unit of compound.

3) Mixtures:

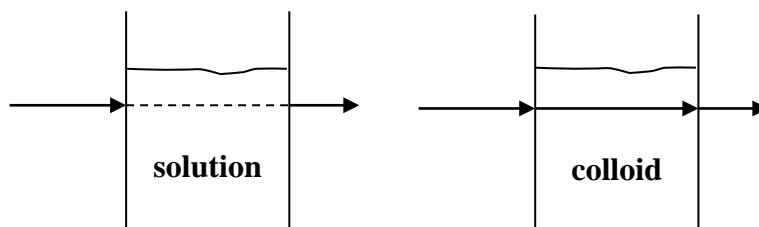
A **mixture** is a form of matter that is composed of two or more different elements, two or more different compounds, or one or more elements and compounds. In a mixture, the individual components usually retain their own individual properties. A mixture is usually physically combined (not bonded) and need NOT be composed in a definite (unique) ratio. A mixture is not a substance because all samples do not necessarily have the same composition. A mixture can be either homogeneous or heterogeneous

a) **Homogeneous mixtures** are those in which all of the components are uniformly distributed. Solutions are the ONLY type of homogeneous mixtures! Filtering cannot separate a solution because the components are too small. A solution will not separate if left standing. A beam of light will not be diffracted (bent) and thus will remain unseen when passed through a liquid or gaseous solution. **Alloys** are also solutions because they are a homogeneous mixture of two or more metals! Alloys are usually not compounds because the elements do not always have to be combined in a unique ratio! Brass, bronze, and steel are examples of alloys.

b) **Heterogeneous mixtures** are those in which the components are not uniformly (evenly) distributed. **Suspensions** and **colloids** are two examples of heterogeneous mixtures.

A suspension is a heterogeneous mixture in which some of the component particles are large enough to be seen by the naked eye and which will separate when left standing.

A colloid (sometimes called a colloidal suspension) is a heterogeneous mixture in which the components are microscopic and will not separate when left standing. A colloid is a mixture much like a solution except that the particles are slightly larger and are NOT homogeneously distributed. A beam of light passing through a **clear** colloid CAN be seen because, although the particles are *too small to be seen*, they are large enough to reflect light. Blood, dusty air, milk, gelatin, whipped cream, colored gems, and marshmallows are all examples of colloids.

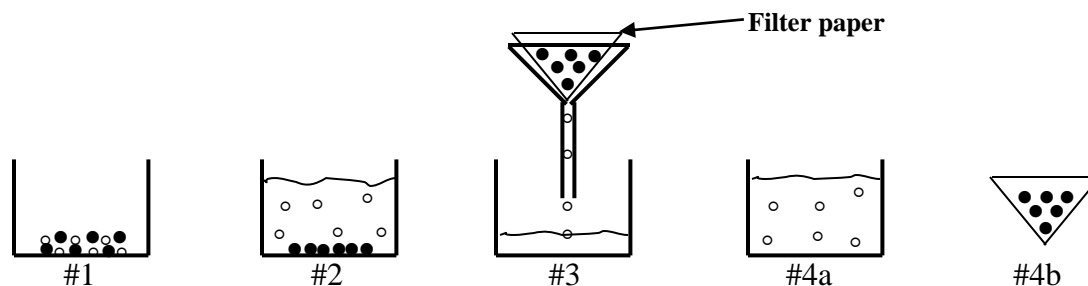


Mixtures can be separated into their individual component parts by various methods. Differences in physical properties such as mass, particle size, molecular polarity, boiling and freezing points, and solubility determine the best method to be used when separating a particular mixture.

1) Filtration:

Filtration is used to separate a water soluble solid from a water insoluble solid (#1). A mixture of the two solids is placed into a beaker and water is added. The water soluble solid dissolves into the solvent (#2). A piece of filter paper is placed in a funnel and the mixture is poured through the funnel where the filter traps the water insoluble solid while allowing the dissolved water soluble solid to pass through with the water (#3). The substances are now separated (4a & 4b)

Examples of Mixtures: (1) sand and $\text{NaCl}_{(\text{aq})}$ (2) $\text{CaCl}_{2(\text{aq})}$ and $\text{CaCO}_{3(\text{s})}$

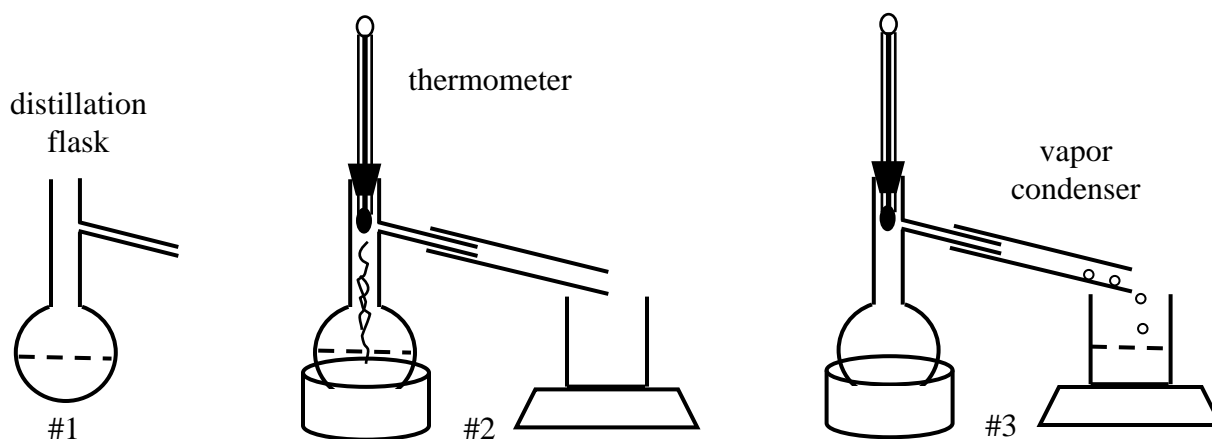


2) Distillation:

Distillation is used to separate a soluble solid solute from a liquid solvent (#1). The solution is boiled. The liquid solvent evaporates leaving the now solid solute behind (#2). The solvent's vapor is collected, cooled and condensed back to a now pure liquid (#3). ***This is how distilled water is made!***

Distillation can also be used to separate a soluble liquid solute from a liquid solvent (#1). The solution is boiled. The liquid with the lower boiling point evaporates leaving the liquid with the higher boiling point behind (#2). The vapor of the lower boiling point liquid is collected, cooled and condensed back to a now pure liquid (#3). This liquid-liquid distillation only works if the two liquids' boiling points are different by at least 5°C . The thermometer is used to know when the first liquid has been separated. As the first liquid is being separated the temperature remains constant. When there is a sudden rise in temperature, this means that the second liquid is now evaporating and it is time to switch the collection beaker. ***This is how distilled liquors are made!***

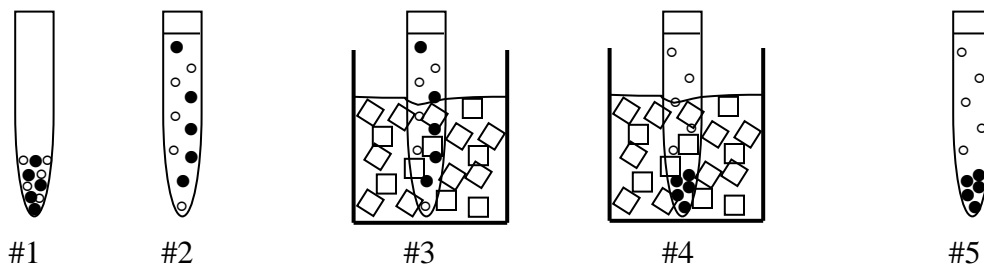
Examples of Mixtures: (1) salt water, (2) alcohol and water [$\text{C}_2\text{H}_5\text{OH}_{(\text{aq})}$]



3) **Crystallization:**

Crystallization is used to separate two solids whose solubility in a cold solvent differs greatly (#1). Both solids are dissolved together in an appropriate liquid solvent (#2). The solution is cooled below the saturation point of the solid of lower solubility (#3). Crystals of the solid of lower solubility begin to form (#4). The mixture can then be filtered (#5).

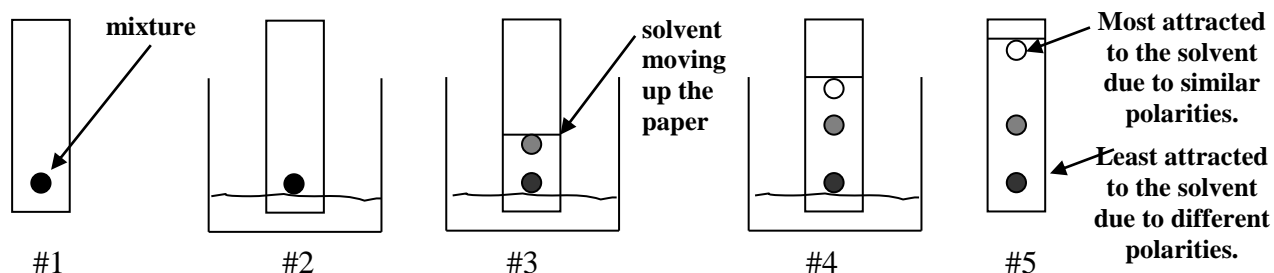
Examples of Mixtures: (1) $\text{KI}_{(\text{aq})}$ & $\text{KNO}_{3(\text{aq})}$

4) **Chromatography:**

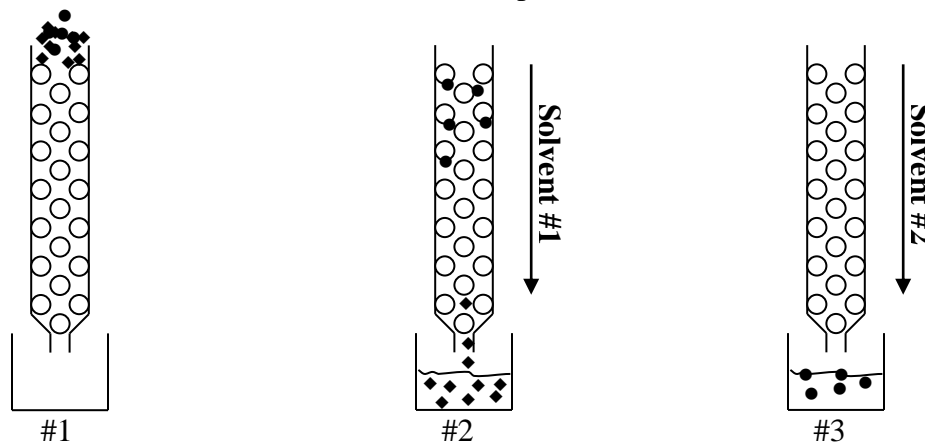
Chromatography is a separation method in which different components of a mixture are attracted differently to a solid surface or to a liquid or a gas solvent. The component of the mixture that has the most similar polarity with the solvent will move right along with the solvent and the component of the mixture that has the least similar polarity with the solvent will be left behind on the paper.

i) **Paper Chromatography:**

A mixture is applied to the bottom of a piece of chromatography paper (#1). The bottom of the paper is suspended into the solvent (#2). The solvent begins to move up the paper and the mixture components begin to separate based on their different polarities (#3, #4, & #5).

ii) **Column Chromatography:**

A mixture is placed at the top of the column which is packed with a special material (#1). A first solvent is sent through the column which readily carries one of the substances in the mixture (but not the other) down the column due to similarities in their polarities (#2). A second solvent is then sent through the column which then readily carries the second substance down the column due to similarities in their polarities (#3).

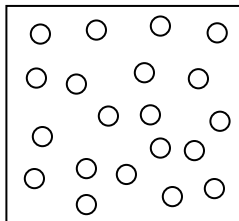


On earth, all forms of matter usually exist in one or more of three forms of matter – **solid, liquid, gas**.

1) **Gases:**

At the macroscopic level, a gas is a form of matter that has (1) no definite shape and (2) no definite volume. A gas will take the shape and volume of the container that it occupies.

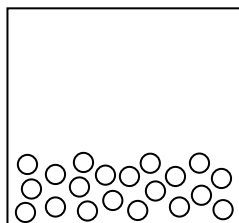
At the submicroscopic level, a gas is composed of tiny, very widely spaced, individual particles (atoms, ions, or molecules). The particles are in a state of continuous, random, rapid, straight-line motion. When particle collisions occur, kinetic energy may be transferred from one particle to another. During this transfer of kinetic energy, none is lost or converted to heat. Pressure results from collisions that the particles make with the walls of the container. ***Gases are highly compressible!***



2) **Liquids:**

At the macroscopic level, a liquid has a definite volume but no definite shape. A liquid takes the shape of their container.

At the submicroscopic level, a liquid is composed of particles (atoms, ions, or molecules) that are tightly packed together. The particles in a liquid are in continuous motion. The particles can vibrate and rotate (spin) around within the group but cannot move in straight lines from place to place due to the close proximity of the particles. ***Liquids are not appreciably compressible!***

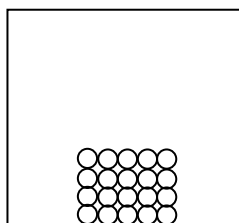


3) **Solids:**

At the macroscopic level, a solid material is one that has a definite shape and a definite volume. A solid's shape is independent of its container.

At the submicroscopic level, a solid is composed of particles (atoms, ions, or molecules) that are tightly packed together. The particles in a solid are in continuous motion. The particles can only vibrate around a fixed position within the group. They can not move around from place to place. ***Solids are not compressible!***

- a) **True solids** are solids in which the particles are arranged in a definite repeating geometric pattern called a **crystal lattice**. True solids melt at a single temperature.
- b) **Amorphous solids** are solids in which the particles are not arranged in a definite repeating geometric pattern. Amorphous solids melt over a temperature range (they get softer and softer before becoming liquid). ***Glass is an amorphous solid!***



D) Intensive and Extensive Properties:

All **properties** can be categorized in various ways. One way is to group properties as either being **intensive** or **extensive**.

1) Intensive Properties:

An **intensive property** is one that is true regardless of the quantity of matter that is present. As such, intensive properties are sometimes called “identifying properties”. Some intensive properties include: density, melting and boiling point, heats of fusion and vaporization, specific heat, phase at room temperature, crystalline shape, and reactivity.

2) Extensive Properties:

An **extensive property** is one that does depend on the quantity of matter that is present. Some extensive properties include: length, mass, volume, and overall shape. Extensive properties alone cannot be used to identify an unknown.

E) Physical and Chemical Properties of Matter:

All properties can also be categorized as either chemical properties or physical properties.

1) Physical Properties:

A **physical property** characterizes any observation that can be made or determined without changing the substance into something new. Some **physical properties** include: phase at room temperature, melting and boiling point, density, solubility, color, texture, and mass.

2) Chemical Properties:

A **chemical property** characterizes how a substance reacts or does not react with another substance to produce something completely new. Some chemical properties include: hydrogen gas will burn (react with oxygen gas) in air, nitrogen gas will NOT burn in air, solid magnesium metal will burn in carbon dioxide, wood will NOT burn in carbon dioxide, cider will ferment, water will NOT ferment, iron (Fe) will oxidize to rust (Fe_3O_4), and stainless steel will NOT oxidize to rust. *Chemical properties are always intensive in nature!*

ALL PROPERTIES ARE FOUND THROUGH OBSERVATION AND EXPERIMENT.

F) Physical and Chemical Changes of Matter:**1) Physical Changes (Physical Reactions):**

A **physical change** is a change resulting in the rearrangement of existing particles in which no new substance is produced. The composition and chemical properties remain the same. Physical reactions may be endothermic (heat is released), exothermic (heat is absorbed), or may not involve energy changes. In exothermic reactions, the reactants release heat to the environment as they change into the products. A physical change involves **NO CHANGES IN COMPOSITION**.

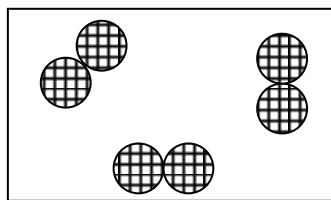
Examples: Dissolving, grinding into powder, melting, boiling, crystallizing and magnetizing.

2) Chemical Changes (Chemical Reactions):

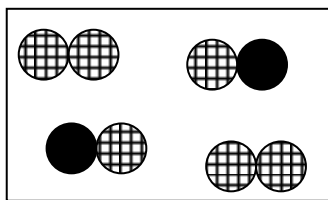
A **chemical change** is a change in which something new is produced. The composition and properties change. Chemical reactions **MUST** be either endothermic or exothermic! In an endothermic reaction, the reactants absorb heat from the environment as they change into the product. A chemical change involves **CHANGES IN COMPOSITION**.

Examples: Burning, fermenting, rusting, polymerizing, photosynthesis, and rotting.

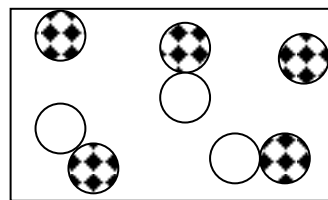
1) Base your answers to the questions below on the following figures.



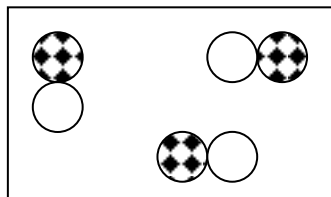
(i)



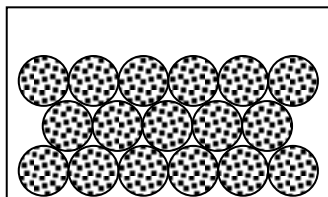
(ii)



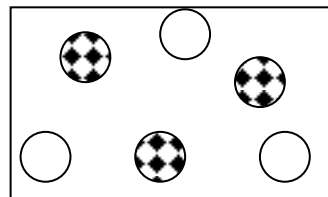
(iii)



(iv)



(v)



(vi)

- Which figures (or figure) represent(s) a pure element? _____
- Which figures (or figure) represent(s) a mixture of two elements? _____
- Which figures (or figure) represent(s) a pure compound? _____
- Which figures (or figure) represent(s) a mixture of an element and a compound? _____

2) Classify each of the following as a pure substance or mixture and if it is homogeneous or heterogeneous.

- air: _____
- cranberry juice: _____
- iodine crystals: _____
- sand: _____

3) Differentiate between a solid, a liquid, and a gas at the macroscopic (what we see) level.

4) Label each of the following as either a physical change or a chemical change.

- iron metal rusts: _____
- ice melts: _____
- aspirin is crushed: _____
- wood decays: _____
- nitroglycerin explodes: _____

A) Measurement:

The system of measurement that is in everyday use in the United States is based on the Imperial (British) system of measurement. This system uses the foot, pound, and second as basic units. One problem with this system is that smaller or larger increments are not standard. There are 3 feet in a yard, but there are 12 inches in a foot! There are 2000 pounds in a ton, but there are 16 ounces in a pound! There are 2 pints in a quart, but there are 4 quarts in a gallon! A second problem with this system is that the pound unit is dependent on the earth's gravitational field, which is not always constant.

1) History:

The **S.I. system of measurement** (a metric system) was first established in France in 1791 as a result of the French Revolution. The adoption of this system by the French revolutionaries was a manifestation of the Age of Reason. They wished to establish a “rational” basis for every aspect of life, and the decimal system satisfied such a scheme. They were also able to rid all measurement from being linked to a monarchy. As it turns out the **S.I. (metric) system** has a further advantage in that the units are absolute since the units are not defined in such a way that local variation would change.

2) S.I. Base Units:

The metric system is sometimes called the “**mks**” system for three of its **base units**; the **meter**, the **kilogram**, and the **second**. A **base unit** is a fundamental unit. In the S.I. system, there are seven **base units**: the meter, the kilogram, the second, the mole, the kelvin, the ampere, and the candela. You will use all of these units this year except for the candela.

- (a) Meter (m): The **meter** is the S.I. base unit of length.
- (b) Kilogram (kg): The **kilogram** is the S.I. base unit of mass.
- (c) Second (s): The **second** is the S.I. base unit of time.
- (d) Kelvin (K): The **kelvin** is the S.I. base unit of temperature.
- (e) Mole (mol): The **mole** is the S.I. base unit of number (of particles of matter).
- (f) Ampere (A): The **ampere** is the S.I. base unit of electric current.
- (g) Candela (cd): The **candela** is the S.I. base unit of luminous intensity (brightness).

3) S.I. Derived Units:

Everything that people know how to measure can be measured using the seven S.I. base units or units derive from them. The following **S.I. derived units** are very important in chemistry.

- (a) Pascal (Pa): The **pascal** is the S.I. derived unit of pressure. $[1 \text{ Pa} = 1 \text{ kg/m}\cdot\text{s}^2]$
- (b) Joule (J): The **joule** is the S.I. derived unit of energy. $[1 \text{ J} = 1 \text{ kg}\cdot\text{m}^2/\text{s}^2]$

4) Non-S.I. Metric Units:

There are some units that are **non-S.I. units**, but which ARE metric units. The following are three important non-S.I. metric units.

- (a) Liter (L): The **liter** is a metric unit for volume. 1 L is equal to 0.001 m^3 .
- (b) Degree Celsius (°C): The **degree Celsius** is a metric unit of temperature. $1.0 \text{ }^\circ\text{C}$ is NOT equal to 1.0 K . However, a **Celsius** temperature is equal to a Kelvin temperature minus 273 ($^\circ\text{C} = \text{K} - 273$).

5) S.I. Prefixes:

S.I. prefixes are prefixes that can be used with S.I. base units, S.I. derived units, and non-S.I. metric units to form new units that are greater than or less than the original units by some multiple or sub-multiple of 10. The following are common **S.I. prefixes**.

kilo – XXX – XXX – “no prefix” – deci – centi – milli – XXX – XXX – micro – XXX – XXX – nano – XXX – XXX – pico

10^3 10^2 10^1 10^0 10^{-1} 10^{-2} 10^{-3} 10^{-4} 10^{-5} 10^{-6} 10^{-7} 10^{-8} 10^{-9} 10^{-10} 10^{-11} 10^{-12}

3 2 1 0 -1 -2 -3 -4 -5 -6 -7 -8 -9 -10 -11 -12

k X X "none" d c m X X μ^* X X n X X p

Thus “kilo” mean thousand, “centi” means hundredth, “micro” means millionth, “nano” means billionth, and “pico” means trillionth.

*This is the Greek letter mu (pronounced “mew”).

B) Dimensional Analysis (the factor-label-method-of-problem-solving):

Dimensional analysis is a powerful systematic tool that is used in solving mathematical problems. The key to using dimensional analysis is the correct use of conversion factors.

A **conversion factor** is fraction (a ratio) whose numerator and denominator are the same quantity expressed in different units. 1.00 inch and 2.54 centimeters are equal quantities. These two equal quantities can be written as the following conversion factors:

$$\frac{1.00 \text{ in}}{2.54 \text{ cm}} \quad \text{or} \quad \frac{2.54 \text{ cm}}{1.00 \text{ in}}$$

Example-1: You can use the first conversion factor above to convert 100.00 centimeters into inches.

$$100.00 \text{ cm} \times \frac{1.00 \text{ in}}{2.54 \text{ cm}} = 39.4 \text{ in}$$

desired unit

given unit

Notice that when the given quantity is multiplied by the conversion factor, the given units “cancel out” and the desired unit remain.

Example-2: You can use the second conversion factor above to convert 70.00 inches (the height of a 5 foot 10 inch person) into centimeters.

$$70.00 \text{ in} \times \frac{2.54 \text{ cm}}{1.00 \text{ in}} = 178 \text{ cm}$$

desired unit

given unit

Notice that when the given quantity is multiplied by the conversion factor, the given units “cancel out” and the desired unit remain.

A general scheme for using dimensional analysis is as follows.

$$\text{quantity in given unit} \times \frac{\text{desired unit}}{\text{given unit}} = \text{quantity in desired unit}$$

conversion factor

Multiple conversion factors may be used in one problem!

Example-3: How many centimeters are in 5 km?

$$5 \text{ km} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{100 \text{ cm}}{1.00 \text{ m}} = 500,000 \text{ cm}$$

Example-4: How many deciliters are in 100 microliters?

$$100 \text{ } \mu\text{L} \times \frac{1.00 \text{ L}}{1,000,000 \text{ } \mu\text{L}} \times \frac{10 \text{ dL}}{1.00 \text{ L}} = 0.001 \text{ dL}$$

Example-5: The Sears tower in Chicago is 1454 ft tall. Calculate its height in meters.

$$1454 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} \times \frac{2.54 \text{ cm}}{1.00 \text{ in}} \times \frac{1.00 \text{ m}}{100 \text{ cm}} = 443 \text{ m}$$

Example-6: The average speed of a nitrogen molecule in air at 25°C is 515 m/s. Convert this speed to miles per hour [1 mi = 1.6093 km].

$$515 \frac{\text{m}}{\text{s}} \times \frac{1 \text{ km}}{1000 \text{ m}} \times \frac{1 \text{ mi}}{1.6093 \text{ km}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{60 \text{ min}}{1 \text{ hr}} = 1150 \frac{\text{mi}}{\text{hr}}$$

Notice that all of the units “cancel out” except for mi/hr!

Conversion factors involving volume

Conversion factors for volume can be tricky. **Remember, everything gets cubed!**

Example-7: What is the volume of 12.00 cubic inch in cubic centimeters?

Since 1.00 in equals 2.54 cm, then $1.00 \text{ in}^3 = 2.54 \text{ cm}^3$...right? **WRONG!**

$$(1.00 \text{ in})^3 = (2.54 \text{ cm})^3 \rightarrow \mathbf{1.00 \text{ in}^3 = 16.4 \text{ cm}^3} \leftarrow \text{the conversion factor}$$

$$12.00 \text{ in}^3 \times \frac{16.4 \text{ cm}^3}{1.00 \text{ in}^3} = 197 \text{ cm}^3$$

Example-8: How many liters are in 1.0 m³? [1.0 L = 1.0 dm³] also [1.0 L = 1000 cm³]

$$1.0 \text{ m}^3 \times \frac{(10 \text{ dm})^3}{(1 \text{ m})^3} \times \frac{1.0 \text{ L}}{1.0 \text{ dm}^3} = 1000 \text{ L}$$

C) Temperature Systems and Temperature Conversions:

In 1714, the first accurate thermometer was made by Gabriel Fahrenheit when he sealed mercury into a glass tube. He set 0 as the lowest temperature he could reach by freezing heavily salted water hoping to avoid negative temperatures. He set 100 as human body temperature. Thus the fahrenheit scale was born. Today, of course, normal human body temperature is 98.6°C

In 1742, Anders Celsius weirdly reset the standard high and low temperature settings as 100 for the freezing point of water and 0 as the boiling point of water! Soon, however, the values were reversed and the celsius scale was born. Actually, Celsius called his scale the centigrade scale and it was called this until 1948 when it was renamed the celsius scale in honor of Celsius.

In 1848, Fahrenheit's idea of a temperature scale with no negative numbers was achieved by William Thomson (Lord Kelvin). Kelvin observed that Charles law, which states that gases lose 1/273 of their volume for every degree below 0°C, implies that at -273°C, the volume would become zero! Kelvin proposed that -273°C was the lowest obtainable temperature now called absolute zero. In honor of Kelvin, -273°C was reset as 0 K and 0°C was reset as 273 K and the kelvin scale was born (notice that the kelvin scale does not use the degree symbol as part of the unit).

Below are six important temperatures that you should know by heart. You probably know many of them already.

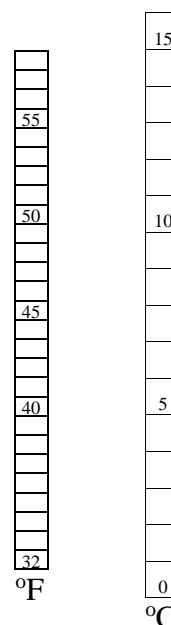
	°F	°C	K
Boiling point of water	212	100	373
Human body temperature	98.6	37	310
Standard room temperature	68	20	293
Freezing point of water	32	0	273
Fahrenheit/Celsius numeric equivalence	-40	-40	233
Absolute zero	-459	-273	0

Know both of the temperature conversion equations below by heart.

$$K = ^\circ C + 273 \quad \text{and} \quad ^\circ F = 1.8^\circ C + 32$$

Notice that a 1°F increase in temperature is a smaller increase than a 1°C increase.

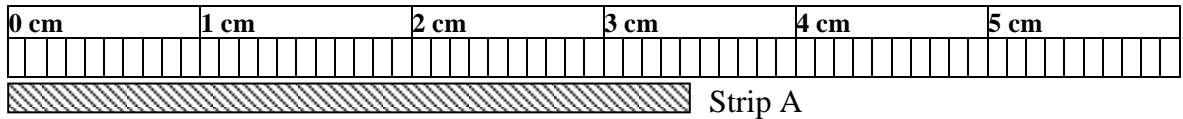
However, a 1.0 K increase in temperature is equal to a 1.0°C increase.



D) Reading and Recording Measurements:

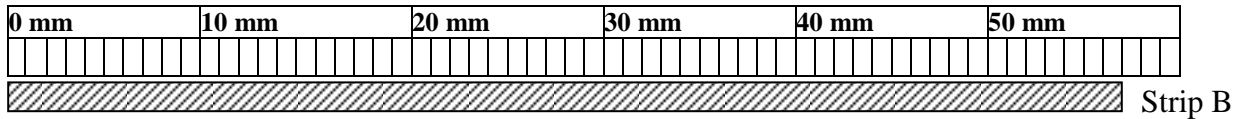
Whenever you use a graduated instrument, you must first determine the unit that the instrument is using, and second, you must determine the smallest increment of that unit on the instrument. In science, graduated instruments are metric in nature (i.e. **each increment is subdivided into 10 smaller increments**). When recording a measurement, you will read and record all of the known digits from the largest to the smallest increments and then you will record an estimated digit that corresponds to an imaginary set of 10 increments between the smallest actual increments.

Ex. (1) What is the length of each of the metal strips below?



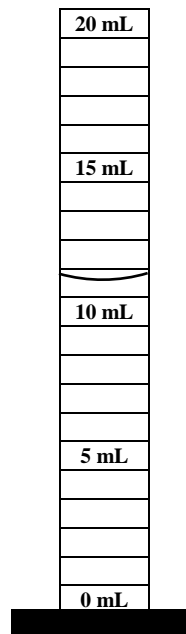
Strip A is _____ cm

Ex. (2) What is the length of each of the metal strips below?



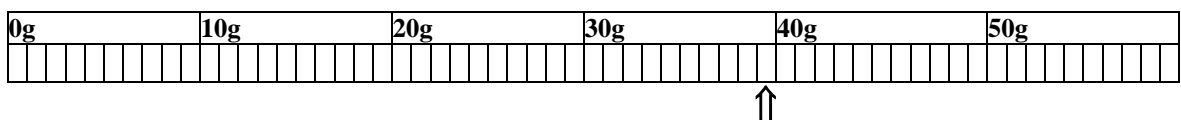
Strip B is _____ mm

Ex. (3) What is the volume of the water in the graduated cylinder below?



The volume of the water is _____ mL

Ex. (4) What is the mass of the object on the balance below?



The mass of the object is _____ g

E) Uncertainty in Measurements:

There are two kinds of numbers in scientific work: *exact numbers* and *inexact numbers*.

Exact numbers can result from counting numbers of objects. For example, there are 18 groups on the Periodic Table. Exact numbers can also be a result of defined values. By definition, an atom of carbon-12 has a mass of 12 amu.

Inexact numbers are values that result from measuring. Since all measurements contain an estimated digit, all measurements contain uncertainty. Scientists try to limit the uncertainty as much as possible but they cannot eliminate it. There are two reasons for uncertainty in measurements.

1) Reasons for Uncertainty:

(a) Instrumental Error:

All measuring instruments have error. The more sensitive and precise the instrument is, the lower the amount of error will be. It is important that instruments be properly maintained and frequently calibrated.

(b) Observer Error:

An instrument is only as good as the person using it! Persons who have more experience and who take more precautions will generally record measurements with less error. Even so, there will be error. This is why measurements are repeated... in order to reduce error. **When discussing error in a report, NEVER talk about observer error! If you do, you will be asked, "Why did you not redo the experiment?"**

2) How to Measure Uncertainty:

In an experiment, it is important to be able to state the level of confidence of one's data. This is done through statistical analysis. Often this analysis is quite involved. However, for this course we will use two simple tests for error.

Accuracy measures how close a measured value is to the accepted value. A test for accuracy uses the equation for Percent (%) Error.

$$\% \text{ Error} = \frac{\text{Experimental Value} - \text{Accepted Value}}{\text{Accepted Value}} \times 100$$

Percent error can be positive or negative! **Positive percent error** means that the experimental value is greater than the accepted value. **Negative percent error** means that the experimental value is less than the accepted value.

Precision measures how close several measured trials are to one another. A **crude** test for precision uses the equation for Percent (%) Range.

$$\% \text{ Range} = \frac{\text{Highest Value} - \text{Lowest Value}}{\text{Average Value}} \times 100$$

In all experiments is desirable to have a low percent error value and a low percent range value.

F) Significant Figures:

When recording a measurement you record all of the known digits and one estimated digit.

These digits are called **significant figures**. It makes no sense to record the volume of a sample of water in a graduated cylinder as 7.5213 mL if the smallest increment on the graduated cylinder is 1.0 mL. Only the digit "7" is truly known and the digit "5" is estimated. The other recorded digits (the "2", the "1", and the "3") would have no significance to the measurement because they would be complete guesswork.

Not all digits in a measurement are significant figures. For example, if the volume of a sample of water in a graduated cylinder is recorded as 7.5 mL, then there are two significant figures in the measurement. If this same measurement were then converted to liters, then the measurement would read 0.0075 L. Since the two measurements are exactly the same, then they should both contain exactly the same number of significant figures – TWO (the "7" and the "5")! The three zeros in 0.0075 L are not significant to the measurement. Of course the two zeros after the decimal point are necessary to hold the empty decimal places, but they are not significant to the measurement.

The following rules are used to determine which digits are significant in a measurement.

- 1) All nonzero digits in a measurement are always significant!
- 2) Any zeros *trapped* or *sandwiched* between two nonzero digits are always significant.
- 3) Any *leading zeros* (zeros to the left of the first nonzero digit) are never significant.
- 4) Any *trailing zeros* (zeros to the right of the last nonzero digit) are ...
 - i) always significant if a decimal point appears in the measurement.
 - ii) never significant if a decimal point does NOT appear in the measurement.
- 5) When a measurement is written in scientific notation, every digit in the first factor is always significant, including any zeros.

Ex. (1) Underline the digits that are significant in each of the following measurements.

- (a) 1020 μ L (b) 0.0027 m (c) 10.0 g (d) 3200. ns

Ex. (2) Underline the digits that are significant in each of the following measurements.

- (a) 36.0°C (b) 98000 Pa (c) 0.01 g (d) 6170 mg

Ex. (3) Underline the digits that are significant in each of the following measurements.

- (a) 35.0°C (b) 35.03°C (c) 10400 mmol (d) 00029 J

Ex. (4) Underline the digits that are significant in each of the following measurements.

- (a) 4.1×10^{-3} g (b) 0.157 lb (c) 30.07 m (d) 0.0100 s

Ex. (5) Underline the digits that are significant in each of the following measurements.

- (a) 775 cal (b) 500. s (c) 0.01010 kg (d) 2.50 cm

Ex. (6) Underline the digits that are significant in each of the following measurements.

- (a) 2.70 g/mL (b) 40 K (c) 007 J (d) 2.500 cm

G) Calculating with Measurements Using the Rules of Significant Figures:

Since **ALL** measurements will contain an estimated digit, then **ALL** measurements will contain some uncertainty (or error). When measurements are added to, subtracted from, multiplied by, or divided by other measurements or numbers, the amount of error in the new measurement is increased. In order to control for this new error, rules for rounding the numeric component of mathematically derived measurements have been established.

1) The Addition and Subtraction Rule: (rounding to a certain number place)

When adding or subtracting a group of measurements, the final answer is rounded so that its last significant figure is in the same number place as the **least most sensitive number place** of the original group of measurements.

$$\text{Ex. (1) } 3.1000\text{g} + 2.67\text{ g} + 1.954\text{ g} = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

$$\text{Ex. (2) } 7.83\text{ mol} - 1.5\text{ mol} = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

$$\text{Ex. (3) } 427\text{ cal} + 173.6\text{ cal} = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

$$\text{Ex. (4) } 8.2\text{ mL} - 1.2\text{ mL} = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

$$\text{Ex. (5) } 1360\text{ mm} + 8300\text{ mm} = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

2) The Multiplication and Division Rule: (rounding to a certain number of significant figures)

When multiplying or dividing a measurement by other measurements, the final answer is rounded so that the number of significant figures is equal to the **least number of significant figures** in the original group of measurements.

$$\text{Ex. (6) } 458.800\text{ cal} \div 4.830\text{ g} = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

$$\text{Ex. (7) } 3.14 \times (2.15\text{ cm})^2 \times 7.00\text{ cm} = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

$$\text{Ex. (8) } 2.50\text{ mm} \times 2.8\text{ mm} = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

WARNING! When you multiply or divide a measurement by a true NUMBER, you round your answer to the same decimal place as the original measurement.

$$\text{Ex. (9) } 3.018\text{ g} \times 25 = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

$$\text{Ex. (10) } 2.13\text{ cm} \div 2 = \underline{\hspace{2cm}} \longrightarrow \underline{\hspace{2cm}}$$

(calculator answer) (final rounded answer)

H) Science and the Scientific Method:

Science is a systematic, self-correcting, ever-evolving way of thinking about the world around us. The triumphs and the mistakes of one generation of scientists become the foundation upon which the next generation of scientists builds their understanding of the universe.

The general approach to doing science is called the **scientific method** which at its core involves the search for truth. The scientific method consists of a set of guidelines which include:

- 1) making observations
- 2) seeking patterns in the observations
- 3) formulating a hypothesis to explain the observations
- 4) designing and conducting experiments to test the hypothesis

A **hypothesis** is a tentative explanation for an observation that is used as a guide in planning future experiments.

The results of many experiments can sometimes be summarized in a short statement or mathematical equation called a **scientific law**. A scientific law does not attempt to explain the behavior of nature but simply to describe it. A scientific law often relates two variables.

Over time, any hypothesis that is continually able to predict and explain facts that have not yet been observed can become a **theory**. In science, a theory is an explanation of the general causes of certain phenomena, with considerable evidence or facts to support it. In science, theories can never be proven absolutely correct, but they can be disproved.

Be aware that, even though science is presented in class as a set of neat and concise ideas, the actual activity of science is anything but neat and concise.

Answers:

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- 2) $3.45(\pm 0.01)$ cm 2) $57.0(\pm 0.1)$ mm 3) $11.6(\pm 0.1)$ mL 4) $39.5(\pm 0.1)$ g

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- | | | | |
|--------------------------------------|-------------------------------|-----------------------|---------------------|
| 1) (a) <u>1020</u> μ L | (b) <u>0.0027</u> m | (c) <u>10.0</u> g | (d) <u>3200.</u> ns |
| 2) (a) <u>36.0</u> $^{\circ}$ C | (b) <u>98000</u> Pa | (c) <u>0.01</u> g | (d) <u>6170</u> mg |
| 3) (a) <u>35.0</u> $^{\circ}$ C | (b) <u>35.03</u> $^{\circ}$ C | (c) <u>10400</u> mmol | (d) <u>00029</u> J |
| 4) (a) <u>4.1</u> $\times 10^{-3}$ g | (b) <u>0.157</u> lb | (c) <u>30.07</u> m | (d) <u>0.0100</u> s |
| 5) (a) <u>775</u> cal | (b) <u>500.</u> s | (c) <u>0.01010</u> kg | (d) <u>2.50</u> cm |
| 6) (a) <u>2.70</u> g/mL | (b) <u>40</u> K | (c) <u>007</u> J | (d) <u>2.500</u> cm |

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- | | | | |
|--|--|--|---------------------------|
| 1) $7.724 \rightarrow 7.72$ g | 2) $6.33 \rightarrow 6.3$ mol | 3) $600.6 \rightarrow 601$ cal | 4) $7 \rightarrow 7.0$ mL |
| 5) $9660 \rightarrow 9700$ mm | 6) $94.98964803 \rightarrow 94.99$ cal/g | 7) $101.60255 \rightarrow 102$ cm ³ | |
| 8) $7 \rightarrow 7.0$ mm ² | 9) $75.45 \rightarrow 75.450$ g | 10) $1.065 \rightarrow 1.07$ cm | |

1) What are the SI units for the following physical quantities?

- a) Mass - _____
- b) Temperature - _____

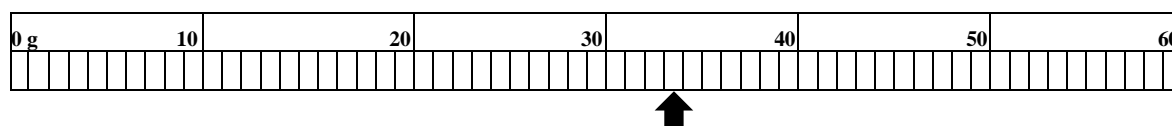
2) What are the metric prefixes, if any, for the following factors?

- a) 10^{-2} - _____
- b) 10^0 - _____
- c) 10^{-9} - _____

3) Underline the digits that are significant in the following measurements.

- a) 0.0320 g
- b) 570 mL
- c) 90.4 °C
- d) 1.60×10^3 m

4) Read and record the mass of a metal ball using the balance to the proper number of significant figures.



The mass of the metal strip is _____ g

5) Make the following temperature conversions.

- a) 25°C to °F
- b) 253 K to °F

6) What is the normal boiling point of water in kelvin? _____ K

7) What is absolute zero in degrees celsius? _____ °C

8) Using the factor-label method, perform the following conversion. The speed of light in a vacuum is 3.00×10^8 m/s. What is the speed of light in miles per hour? [2.54 cm = 1.00 in; 5280 ft = 1.00 mi;]

A) The Composition and Structure of the Atom:**1) The Atomic Theory of Matter:**

In the first decade of the nineteenth century, John Dalton proposed the first atomic theory that was based on experimental evidence. Dalton's atomic theory involved the following postulates:

- (1) Elements are composed of tiny particles called atoms.
- (2) Atoms of the same element are exactly alike and have the same set of properties, but the atoms of one element are different from the atoms of all other elements.
- (3) Atoms of an element cannot be created, destroyed, or changed into atoms of another element.
- (4) Compounds are formed when atoms of two or more elements are chemically combined in a fixed small whole number ratio.

What evidence did Dalton use to develop his theory?

- (1) Dalton knew that compounds were always composed of the same elements in the same proportion by mass. ***This is the Law of Definite (or Constant) Composition.***

Water was always composed of 11.1% hydrogen and 88.9% oxygen.

Carbon dioxide was always composed of 27.3% carbon and 72.7% oxygen.

Mercury (II) sulfide was always composed of 86.2% mercury and 13.8% sulfur.

Dinitrogen monoxide was always composed of 63.6% nitrogen and 36.4% oxygen.

Dalton concluded that the presence of discrete particles of matter (atoms) was the only explanation for the constancy of the composition by mass for compounds because atoms always combined whole and in definite ratios in a compound.

- (2) Dalton knew that, in reactions, the total mass of all substances did not change. The total mass of the products was equal to the total mass of the reactants. ***This is the Law of Conservation of Mass (or Matter).***

When 10.0 g of diamond (carbon) was burned in 26.6 g of oxygen gas, 36.6 g of carbon dioxide was produced (experiment conducted by Antoine Lavoisier).

When 10.00 g of mercury (II) oxide was decomposed, 9.26 g of mercury and 0.74 g of oxygen gas was produced.

Dalton concluded that the presence of discrete particles of matter (atoms) was the only explanation for the conservation of mass during reactions because if you started and ended with the same total mass then the masses of the basic types of matter must be the same.

- (3) Dalton knew that when elements formed more than one compound and when the same amount of one element (element A) was present in all compounds, the amounts of the second element (element B) in each compound would always be found in a small whole number ratio to each other. ***This is the Law of Multiple Proportions.***

*Nitrogen and oxygen can form N_2O - dinitrogen monoxide (5.00 g nitrogen and 2.86 g oxygen), NO - nitrogen monoxide (5.00 g nitrogen and 5.71 g oxygen), and NO_2 - nitrogen dioxide (5.00 g nitrogen and 11.44 g oxygen). With nitrogen constant at 5.00 g in all three compounds, the small whole number ratio of the oxygen in each compound, 2.86 g to 5.71 g to 11.44 g, is **1 to 2 to 4!***

Dalton concluded that the presence discrete particles of matter (atoms) was the only explanation for this phenomenon. **Matter must exist as small whole particles!**

2) The Discovery of Atomic Structure:

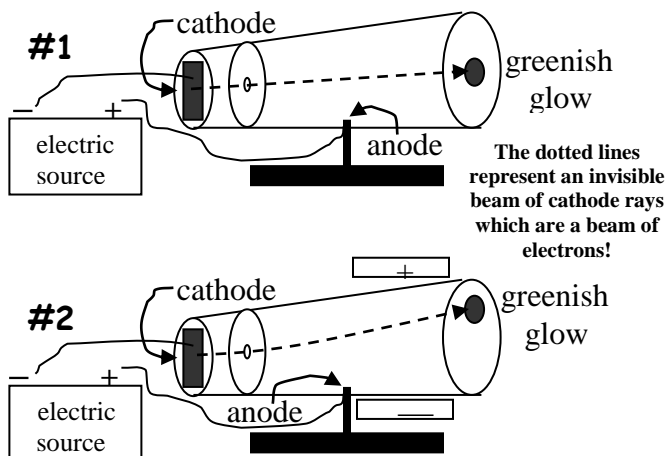
Over time, additional experimental evidence led to the modification of Dalton's theory. The following experiments are some that helped to further develop our knowledge of the atom.

(a) The Cathode Ray Tube Experiments (J.J. Thomson - 1897)

- Found that elements possessed small negatively charged particles (electrons).
- Found that the mass-to-charge ratio of these electrons was 5.68×10^{-9} g/coulomb.

J.J. Thomson placed a barrier with a small opening in front of the cathode (**figure #1**). This helped to show that the cathode rays traveled in a straight beam. He then passed the beam of cathode rays between oppositely charged electric plates. The beam was always deflected toward the positive plate (**figure #2**). This showed that cathode rays were negatively charged. Thomson concluded that cathode rays consisted of a beam of negatively charged particles - **electrons!**

Thomson also found that the mass-to-charge ratio of electrons was 5.68×10^{-9} g/C

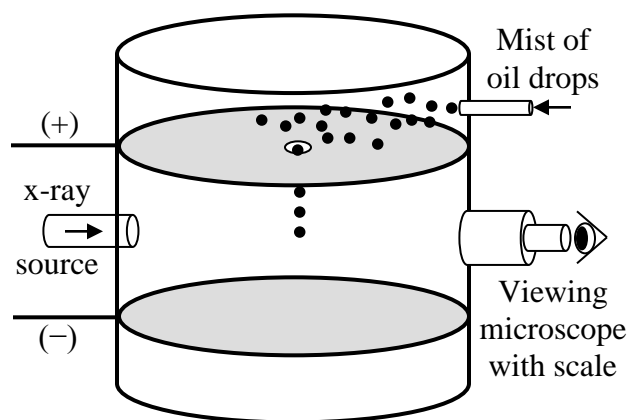


(b) The Oil Drop Experiment (Robert Millikan - 1909)

- Found that the charge of an electron was 1.60×10^{-19} coulombs.
- Using the mass-to-charge ratio of an electron, found that its mass was 9.10×10^{-28} g.

Millikan injected oil droplets into the chamber which acquired extra electrons due to the x-rays. He then adjusted the voltage in order to suspend an oil drop between charge plates. Viewing the drop, he could determine its volume. Using the oil's density he determined the drop's mass. Through complex calculations **he was able to determine the electron's charge to be 1.60×10^{-19} C.**

Using the electron's mass-to-charge ratio, **he found that its mass was 9.10×10^{-28} g**

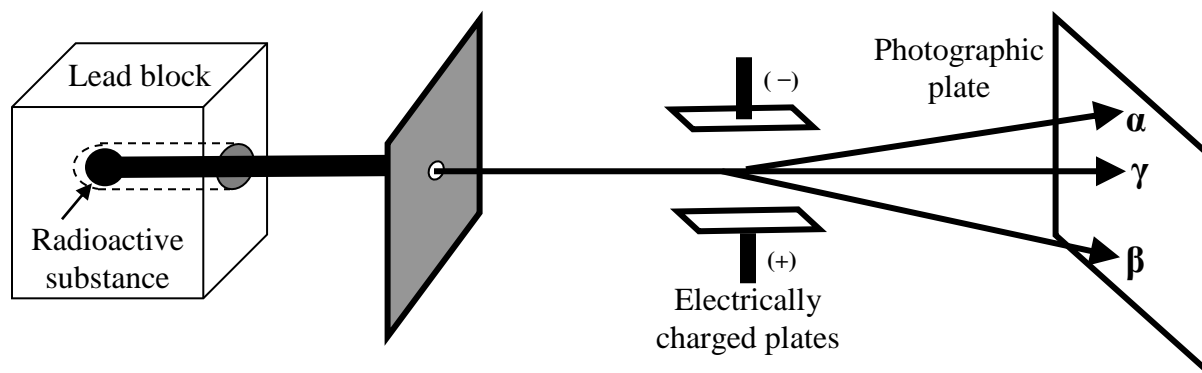


(c) Radioactivity Experiments (Ernest Rutherford et al. - 1887 to 1903)

- Three types of radiation are found. Rutherford names them alpha, beta, and gamma.

Particle	Alpha (α).	Beta (β).	Gamma (γ).
Relative Mass	4 amu	≈ 0 (0.00055) amu*	0 amu
Relative Charge	+2	-1	0 (neutral)

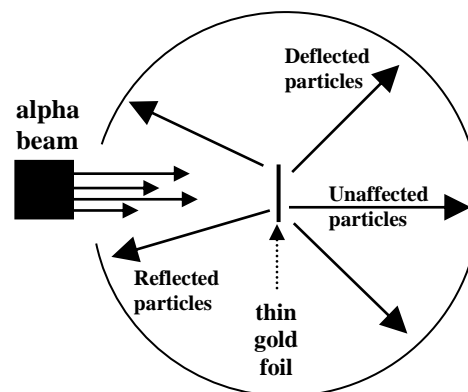
***you only need memorize the relative mass of an electron as 0 amu.**



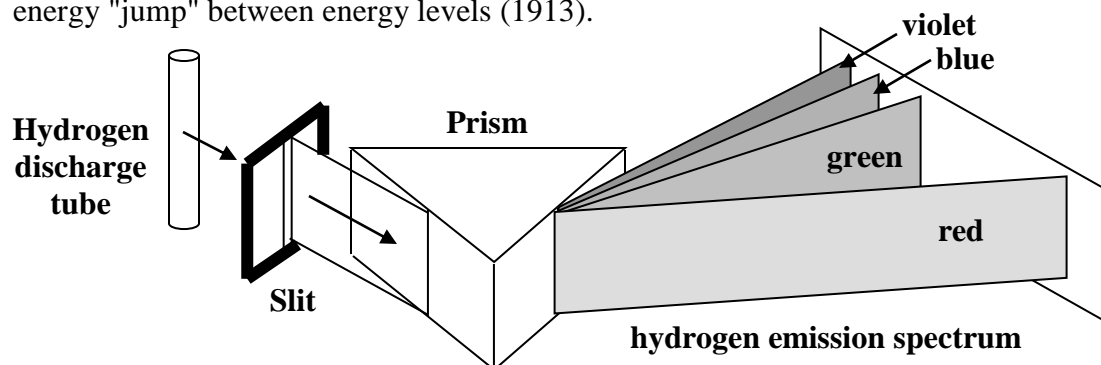
(d) The Gold Foil Experiment (Ernest Rutherford – 1911 to 1913)

- Found that an atom's volume is mostly empty space.
- Found that most of an atom's mass is contained in an extremely small positive charged central volume called the nucleus.
- The proton is discovered (1914)

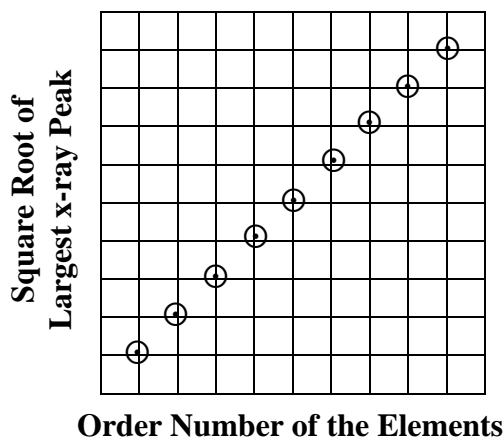
To Rutherford's amazement, about 1 alpha particle in 8,000 to 10,000 seemed to bounce back from the gold foil. Almost all the alpha particles passed right through the gold foil. Less than 1% of the alpha particles were deflected from their straight-line course as they passed through the gold foil.

(e) The Visible Light Emission Spectra Experiments (Johann Balmer et al.- 1859 to 1885)

- Found that excited elements emit a unique set of spectral lines as seen in a spectroscope.
- Neils Bohr determined that each line in the **hydrogen spectrum** represented an electron energy "jump" between energy levels (1913).

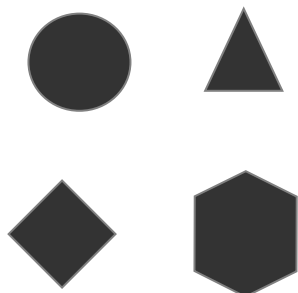
(f) The X-ray Emission Spectra Experiment (Henry Moseley- 1913)

- Found that a plot of the square root of the elements' X-ray emission frequency versus the "order number" of the elements was a perfect straight line.
- Moseley's "order number" is known today as the atomic number (Z).

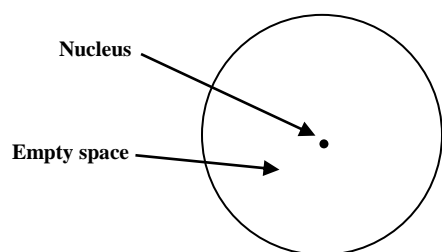
(g) The Alpha Bombardment of Beryllium Experiment (James Chadwick - 1932)

- Found that a neutral particle, similar in mass to a proton, with high penetrating power was ejected from beryllium after being bombarded with alpha particles.
- The neutron was discovered which had been predicted 12 years earlier.

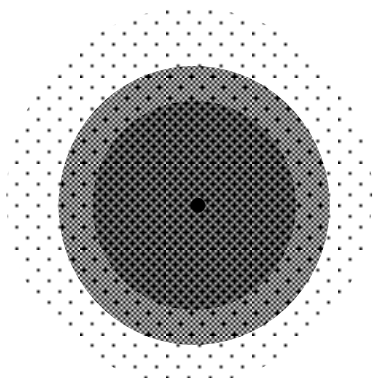
3) The Historical Development of Modern Atomic Theory:



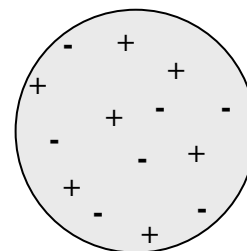
a) Dalton's model of the atom proposed that an atom was a small, solid, indivisible, and indestructible particle and that atoms of different elements were of different sizes and shapes.



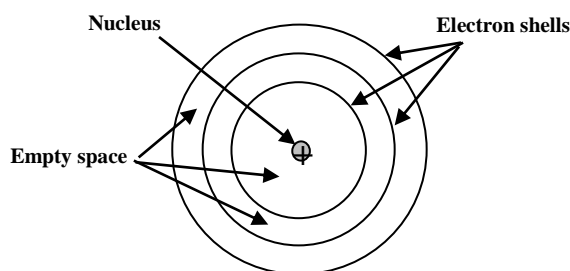
c) Rutherford's model of the atom suggested that an atom's volume is mostly empty space. It has a small (one 10000th the volume of the atom itself), but massive (almost 100% of an atom's mass) nucleus that carries a positive charge due to the presence of protons. The electrons were thought to be at the outer edge of the atom. His model is sometimes referred to as the "empty space" model.



e) The modern model of the atom consists of a diffuse "cloud" of rapidly moving negatively charged electrons occupies nearly all of the atomic volume and surrounds the tiny central nucleus. This model is referred to as the *wave-mechanical* model, the *electron-cloud* model, the *charge-cloud* model, or the *quantum-mechanical* model.



b) Thomson's model of the atom was a positively charged solid that was studded with negatively charged sub-particles. His model is sometimes referred to as the "*plum pudding*" model.



d) Bohr's model of the atom was similar to Rutherford's model except that the electrons orbited the nucleus in discrete energy levels. His model is sometimes referred to as the "*planetary*" model

4) **Subatomic Particles:**

Today we know that atoms are composed of protons, neutrons, and electrons. Study the chart below and know, **by heart**, the information.

Particle	Symbol	Relative mass	Relative Charge	Location
proton	${}^1_1\text{H}$ or p	1 amu	+1	nucleus
neutron	${}^1_0\text{n}$ or n	1 amu	0 (neutral)	nucleus
electron	${}^0_{-1}\text{e}$ or e^-	≈ 0 (0.00055) amu*	-1	outside nucleus

*you only need memorize the relative mass of an electron as 0 amu.

B) Isotopes:

Isotopes are atoms with the same number of protons and different numbers of neutrons.

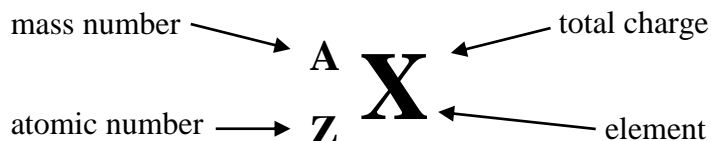
The number of protons is called the **atomic number (Z)**. The number of protons and neutrons is called the **mass number (A)**.

Thus, **isotopes** can also be defined as atoms with the same atomic number and different mass numbers.

Atoms with the same number of protons are all the same element. Atoms with different mass numbers have different masses.

Thus, **isotopes** can also be defined as atoms of the same element with different masses.

Isotopes can be written using their **nuclide symbol**.



Neutral atoms have the same number of electrons as protons. In fact, the term **atom** alone implies a neutral particle.

- (1) Neutral carbon-14 has 6 protons, 6 electrons, and 8 neutrons and is written as ${}^{14}_6\text{C}$.
- (2) Neutral uranium-235 has 92 protons, 92 electrons, and 143 neutrons and is written as ${}^{235}_{92}\text{U}$.
- (3) A 1+ ion of hydrogen-3 (tritium) has 1 proton, 0 electrons, and 2 neutrons and is written as ${}^3_1\text{H}^{1+}$.
- (4) A 1- ion of chlorine-35 has 17 protons, 18 electrons, and 18 neutrons and is written as ${}^{35}_{17}\text{Cl}^{1-}$.
- (5) ${}^{25}\text{Mg}$ has 12 protons, 12 electrons, and 13 neutrons (the protons are known by the element).
- (6) ${}^{19}\text{F}^{1-}$ has 9 protons, 10 electrons, and 10 neutrons (the protons are known by the element).
- (7) ${}^{27}\text{Al}^{3+}$ has 13 protons, 10 electrons, and 14 neutrons (the protons are known by the element).
- (8) ${}^{223}\text{Fr}$ has 87 protons, 87 electrons, and 136 neutrons (the protons are known by the element).
- (9) Pb-206 has 82 protons, 82 electrons, and 124 neutrons (the protons are known by the element).
- (10) Tc-99 has 43 protons, 43 electrons, and 56 neutrons (the protons are known by the element).

C) The Atomic Mass Scale:

Although the existence of protons, neutrons, and electrons were unknown to early scientists, they were aware that atoms of different elements did have different masses. They knew that for every 100.0 g of water, there was 11.1 g of hydrogen and 88.9 g of oxygen. Thus, water contained $88.9/11.1 = 8$ times as much oxygen, by mass, as hydrogen. Using hydrogen as the mass standard and arbitrarily given the mass of 1 (no units), early scientists believed that the relative mass of oxygen was 8! Once it was known that water contained two hydrogen atoms for each oxygen atom, the relative mass of oxygen was changed to 16.

TODAY, the masses of atoms are measured using the **atomic mass unit** (amu).

Carbon-12 as the standard (H-1 and then O-16 were the standards at one time) and C-12 is assigned the mass of **exactly** 12 amu ($1 \text{ amu} = 1.661 \times 10^{-24} \text{ g}$).

Thus the atomic mass unit (1 amu) is defined as **exactly 1/12 the mass of C-12**.

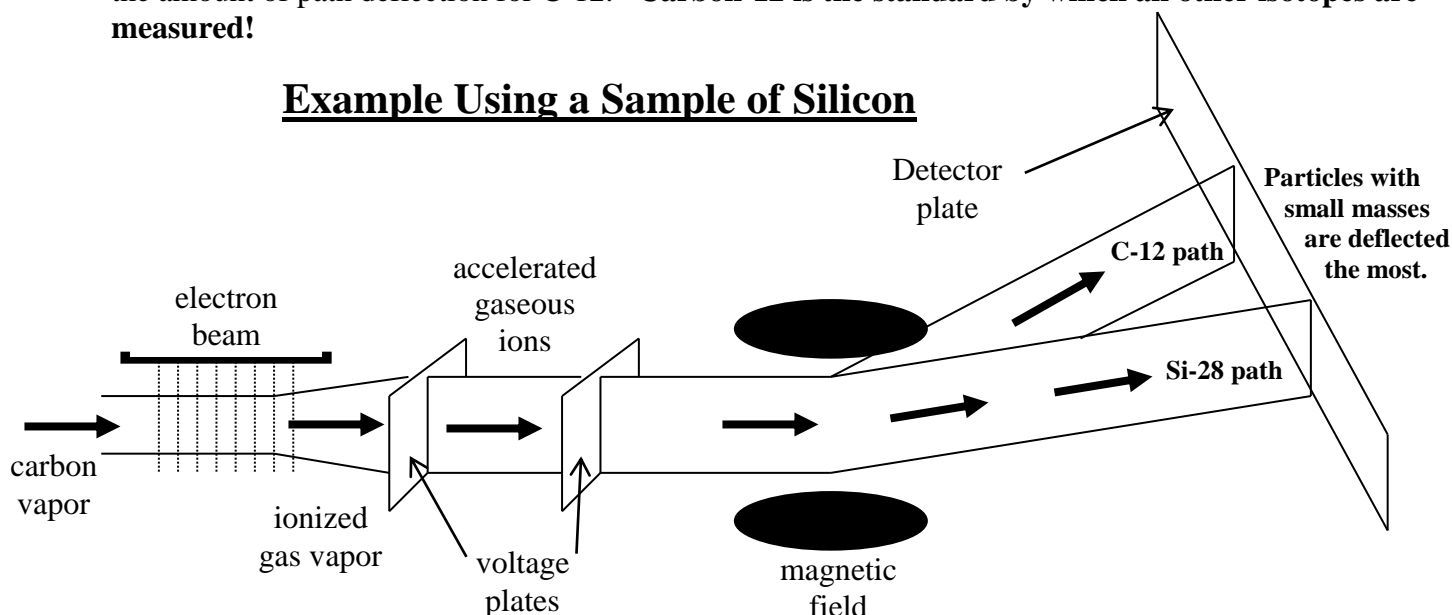
In honor of John Dalton, the atomic mass unit has been renamed the **dalton (D)**. However, this unit has not caught on yet and so, in this course, we will continue to use the **atomic mass unit** (amu).

D) The Mass Spectrometer:

The most accurate method currently available for comparing the masses of atoms or molecules involves the use of an instrument called the *mass spectrometer*.

- In this instrument, atoms or molecules are vaporized and passed into a beam of high-speed electrons which knock electrons off the atoms or molecules making them positive ions.
- These ions are focused into a narrow beam and then, accelerated by an electric field. They are then sent through an applied magnetic field.
- Since moving ions create their own magnetic field, an interaction occurs between the applied magnetic field and the magnetic field of the moving ions which then changes the path of the ions. The lighter the ions are, the greater their path deflections (changes) will be.
- The ions are then collected on a detector plate which can also determine their abundance. *The mass spectrometer then produces a printout which is read and analyzed.*
- The amount of path deflection gives a very accurate value for the relative mass when compared with the amount of path deflection for C-12. **Carbon-12 is the standard by which all other isotopes are measured!**

Example Using a Sample of Silicon



Example-1: When Si-28 and C-12 are analyzed in a mass spectrometer, the ratio of their path deflections is found to be 2.331411.

Thus, the ratio of their masses is found to be ... $\frac{\text{Mass } ^{28}\text{Si}}{\text{Mass } ^{12}\text{C}} = 2.331411$

Since the mass of ^{12}C is exactly 12 atomic mass units (amu) by definition, then

$$\text{Mass } ^{28}\text{Si} = (2.331411)(12.000000) = \underline{\underline{27.97693 \text{ amu}}}$$

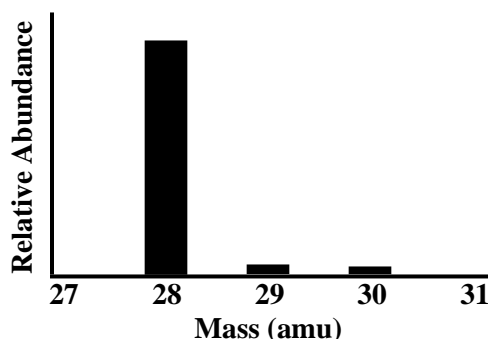
Example-2: When He-4 and C-12 are analyzed in a mass spectrometer, the ratio of their path deflections is found to be 0.3336.

Thus, the ratio of their masses is found to be ... $\frac{\text{Mass } ^4\text{He}}{\text{Mass } ^{12}\text{C}} = 0.3336$

Since the mass of ^{12}C is exactly 12 atomic mass units (amu) by definition, then

$$\text{Mass } ^4\text{He} = (0.3336)(12.0000) = \underline{\underline{4.003 \text{ amu}}}$$

The printout of the mass spectrum for silicon is show below. It shows that silicon has three isotopes whose abundances are: 92.23% Si-28, 4.68% Si-29, and 3.09% Si-30.



E) The Atomic Masses of the Elements:

Most elements occur in nature as a mixture of isotopes. Each isotope has its own atomic mass. Why then does the Periodic Table list only one mass for each element? The atomic masses of the elements that are listed on the Periodic Table are a **weighted average** of the atomic masses of the different naturally occurring isotopes for each of the elements.

An element's **weighted average** can be calculated by multiplying the fractional abundance for each isotope by its mass and then by summing the resulting products.

Example-1: Naturally occurring silicon is composed of 92.23% Si-28, 4.68% Si-29, and 3.09% Si-30. The mass of Si-28 is 27.97693 amu, the mass of Si-29 is 28.97649 amu, and the mass of Si-30 is 29.97377 amu.

$$\text{Weighted average} = (0.9223)(27.97693 \text{ amu}) + (0.0468)(28.97649 \text{ amu}) + (0.0309)(29.97377 \text{ amu})$$

$$\text{Weighted average} = (25.80 \text{ amu}) + (1.36 \text{ amu}) + (0.926 \text{ amu})$$

Weighted average = 28.09 amu

Example-2: Boron has two naturally occurring isotopes: B-10 and B-11. The atomic mass of B-10 is 10.0129 amu and the atomic mass of B-11 is 11.0093 amu. The average atomic mass for the element boron is 10.81 amu. From this data, calculate the percent abundances for B-10 and B-11.

$$10.81 = (X)(10.0129 \text{ amu}) + (1-X)(11.0093 \text{ amu}) \quad \text{if } X = \text{abundance B-10, then } 1-X = \text{abundance B-11}$$

$$X = 0.20$$

20% B-10 and 80% B-11

Average atomic mass is known as **atomic weight** or more recently as **atomic mass**. The units for atomic weight (atomic mass) are “amu” for individual atoms or “g/mole” for a mole of atoms.

Gram atomic weight, gram atomic mass, gram formula weight, gram formula mass, and mole mass are other names for atomic weight when using units of gram per mole.

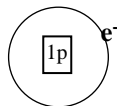
The masses of the elements can be found on the Periodic Table

- 1) Which of the following points was not part of Dalton's atomic theory?
 - (1) Elements are composed of tiny particles called atoms.
 - (2) Atoms of the same element have the same set of properties but may have different masses. The atoms of one element have different properties from the atoms of all other elements.
 - (3) Atoms of an element cannot be created, destroyed, or changed into atoms of another element.
 - (4) Compounds are formed when atoms of two or more elements are chemically combined in a fixed small whole number ratio.
- 2) What experiment was able to determine the existence of electrons?
 - (1) Cathode Ray Tube Experiment
 - (2) Oil Drop Experiment
 - (3) Gold Foil Experiment
 - (4) X-ray Emission Spectra Experiment
- 3) What experiment was able to determine the mass-to-charge ratio of an electron?
 - (1) Cathode Ray Tube Experiment
 - (2) Oil Drop Experiment
 - (3) Gold Foil Experiment
 - (4) X-ray Emission Spectra Experiment
- 4) What experiment was able to determine the charge and then the mass of an electron?
 - (1) Cathode Ray Tube Experiment
 - (2) Oil Drop Experiment
 - (3) Gold Foil Experiment
 - (4) X-ray Emission Spectra Experiment
- 5) What experiment was able to determine the atomic number of an element?
 - (1) Cathode Ray Tube Experiment
 - (2) Oil Drop Experiment
 - (3) Gold Foil Experiment
 - (4) X-Ray Emission Spectra Experiment
- 6) What was the first model of the atom that included electrons?
 - (1) Thomson's model of the atom
 - (2) Rutherford's model of the atom
 - (3) Bohr's model of the atom
 - (4) Wave-Mechanical model of the atom
- 7) What was the first model of the atom that included a nucleus?
 - (1) Thomson's model of the atom
 - (2) Rutherford's model of the atom
 - (3) Bohr's model of the atom
 - (4) Wave-Mechanical model of the atom
- 8) What was the first model of the atom that included energy levels?
 - (1) Thomson's model of the atom
 - (2) Rutherford's model of the atom
 - (3) Bohr's model of the atom
 - (4) Wave-Mechanical model of the atom
- 9) What was the first model of the atom that included electron orbitals?
 - (1) Thomson's model of the atom
 - (2) Rutherford's model of the atom
 - (3) Bohr's model of the atom
 - (4) Wave-Mechanical model of the atom
- 10) What was the first model of the atom that explained the emission spectra of the elements?
 - (1) Thomson's model of the atom
 - (2) Rutherford's model of the atom
 - (3) Bohr's model of the atom
 - (4) Wave-Mechanical model of the atom
- 11) What was the first model of the atom that described an atom's volume as mostly empty space?
 - (1) Thomson's model of the atom
 - (2) Rutherford's model of the atom
 - (3) Bohr's model of the atom
 - (4) Wave-Mechanical model of the atom

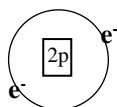
A) The Bohr Model and Electron Energy Shells:

In first year chemistry, you learned that electrons in the Bohr model are found outside the nucleus in various quantized **energy shells**. The maximum number of electrons that a particular energy shell can hold is equal to $2n^2$ where n is equal to the energy shell's number. Thus the first energy shell can hold up to 2 electrons [$2(1)^2 = 2$]. The second energy shell can hold up to 8 electrons [$2(2)^2 = 8$]. The third energy shell can hold up to 18 electrons [$2(3)^2 = 18$]. The fourth energy shell can hold up to 32 electrons [$2(4)^2 = 32$]. The fifth energy shell can hold up to 50 electrons [$2(5)^2 = 50$], etcetera.

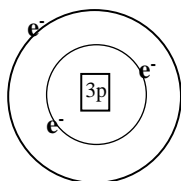
From the Bohr models below, the element's **shell electron configuration** can easily be determined.



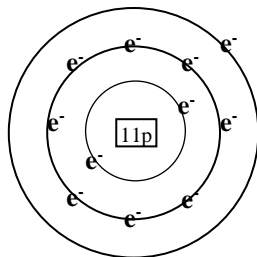
In a hydrogen atom, there is only one electron which is found in the first energy shell (the one closest to the nucleus). The **shell electron configuration** for a hydrogen atom is: **1**



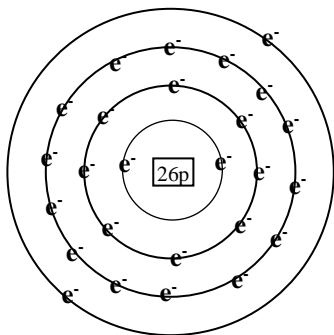
In a helium atom, there are two electrons which are found in the first energy shell. The **shell electron configuration** for a helium atom is: **2**



In a lithium atom, there are three electrons. The first two electrons are found in the first energy shell, and the third electron is found in the second energy shell. The **shell electron configuration** for a lithium atom is: **2-1**



In a sodium atom, there are eleven electrons. The first two electrons are found in the first energy shell, the third through the tenth electrons are found in the second energy shell, and the eleventh electron is found in the third energy shell. The **shell electron configuration** for a sodium atom is: **2-8-1**



In an iron atom, there are twenty six electrons. The first two electrons are found in the first energy shell, the third through the tenth electrons are found in the second energy shell, the eleventh through the twenty fourth electrons are found in the third energy shell, and the twenty fifth and the twenty sixth electrons are found in the fourth energy shell. The **shell electron configuration** for an iron atom is: **2-8-14-2**

Using the shell electron configuration for an element and the element's atomic number (the number of protons), you can easily construct a simple Bohr model for an atom of that element.

In addition, you learned that the **period** in which an element is located on the Periodic Table is equal to the number of occupied energy shells in an atom of that element. Hydrogen is found in period 1 on the Periodic Table and a hydrogen atom has 1 occupied energy shell. Helium is found in period 1 on the Periodic Table and a helium atom has 1 occupied energy shell. Lithium is found in period 2 on the Periodic Table and a lithium atom has 2 occupied energy shells. Sodium is found in period 3 on the Periodic Table and a sodium atom has 3 occupied energy shells. Iron is found in period 4 on the Periodic Table and an iron atom has 4 occupied energy shells.

If you look back at the Bohr model for **iron**, you will notice that fourth energy shell contains two electrons even though the third energy shell, with only fourteen electrons, is not yet filled (the third energy shell can hold up to 18 electrons). ***Why is this so, since electrons in an atom are supposed to occupy the lowest-energy shells available?*** Why isn't iron's shell electron configuration, **2-8-16**, since the third energy shell can hold up to 18 electrons?

As it turns out, the electrons in iron ARE occupying the lowest available energy levels. It's just that the shell electron configuration model **2-8-14-2** does not give enough information about iron's complete electronic structure!

B) Principal Energy Levels and Energy Sublevels:

First, it turns out that energy shells are split into **sublevels!!!**

Thus, what we have been referring to as "energy shells" must now be called **principal energy levels (PEL's)**. The previously called 1st energy shell is now called the **1st PEL**, the previously called 2nd energy shell is now called the **2nd PEL**, the previously called 3rd energy shell is now called the **3rd PEL**, the previously called 4th energy shell is now called the **4th PEL**, and the previously called 5th energy shell is now called the **5th PEL**, etcetera.

The number of sublevels found within a particular PEL is equal to the number of the principal energy level. Thus the 1st PEL is "split" into only 1 sublevel. This is to say that the 1st PEL has only 1 level of energy. Thus for the 1st PEL, and only for the 1st PEL, the sublevel and the PEL itself are the same. The 2nd PEL is split into 2 sublevels which mean that the 2nd PEL has two levels of energy within it. The 3rd PEL is split into 3 sublevels which mean that the 3rd PEL has three levels of energy within it. The 4th PEL is split into 4 sublevels. The 5th PEL is split into 5 sublevels, etcetera.

We designate a particular principal energy level using numbers. The lowest energy PEL (the one closest to the nucleus) is designated the number one - the 1st PEL. The next higher energy PEL is designated the number two - the 2nd PEL. The next higher energy PEL is designated the number three - the 3rd PEL. The next higher energy PEL is designated the number four - the 4th PEL. The next higher energy PEL is designated the number five - the 5th PEL, etcetera.

We designate a particular sublevel in a PEL using letters. The lowest sublevel within a PEL is always designated the letter "s". Thus the lowest (and only) sublevel in the 1st PEL is designated "1s". The lowest sublevel in the 2nd PEL is designated "2s". The lowest sublevel in the 3rd PEL is designated "3s". The lowest sublevel in the 4th PEL is designated "4s". The lowest sublevel in the 5th PEL is designated "5s", etcetera.

The next higher and second sublevel within a PEL is always designated the letter "p". The 1st PEL does NOT have a p-sublevel since the 1st PEL is only "split" into 1 sublevel which is the s-sublevel! The 2nd PEL's second sublevel is designated "2p". The 3rd PEL's second sublevel is designated "3p". The 4th PEL's second sublevel is designated "4p". The 5th PEL's second sublevel is designated "5p", etcetera.

The next higher and third sublevel within a PEL is always designated the letter "d". The 1st PEL and the 2nd PEL do NOT have a d-sublevel since the 1st PEL is only "split" into 1 sublevel which is the s-sublevel and the 2nd PEL is only split into 2 sublevels, the s-sublevel and the p-sublevel! The 3rd PEL's third sublevel is designated "3d". The 4th PEL's third sublevel is designated "4d". The 5th PEL's third sublevel is designated "5d", etcetera.

The next higher and fourth sublevel within a PEL is always designated the letter "f". The 1st PEL, the 2nd PEL, and the 3rd PEL do NOT have an f-sublevel since the 1st PEL is only "split" into 1 sublevel which is the s-sublevel, the 2nd PEL is only split into 2 sublevels, the s-sublevel and the p-sublevel, and the 3rd PEL is only split into 3 sublevels, the s-sublevel, the p-sublevel, and the d-sublevel! The 4th PEL's fourth sublevel is designated "4f". The 5th PEL's fourth sublevel is designated "5f", etcetera.

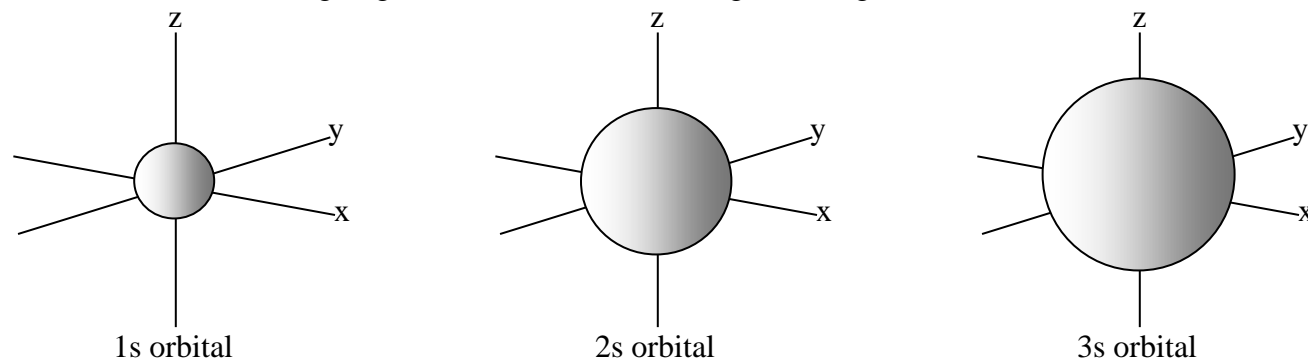
The letters used to designate the first four sublevels in a PEL (s, p, d, f) seem a bit arbitrary. The letters stand for the words “**sharp**”, “**principal**”, “**diffuse**”, and “**fundamental**” which were terms used by scientists to describe the lines in an element’s bright-line emission spectrum. When the field of quantum mechanics was being developed, scientists used the first letter of these four terms to designate the first four sublevels in a PEL (*I don’t know why*). Any sublevel after the “f” sublevel simply goes in alphabetical order. Thus, in increasing order of energy, the sublevels are: **s, p, d, f, g, h, i**, etcetera.

C) Electron Orbitals:

It turns out that within each sublevel there are a specified number of orbitals!!! **Recall, that an orbital is a probability location for an electron in an atom.** All s-sublevels contain one orbital. This means that each orbital in any s-sublevel is that entire s-sublevel! All p-sublevels contain three orbitals. This means that the three orbitals in a p-sublevel are the p-sublevel. All d-sublevels contain five orbitals. This means that the five orbitals in a d-sublevel are the d-sublevel. All f-sublevels contain seven orbitals. This means that the seven orbitals in an f-sublevel are the f-sublevel. Notice that the number of orbitals in successive sublevels increases by two. Thus the g-sublevel contains nine orbitals, the h-sublevel contains eleven orbitals, and the i-sublevel contains thirteen orbitals. And all of these orbitals are potential probability locations for electrons!!!

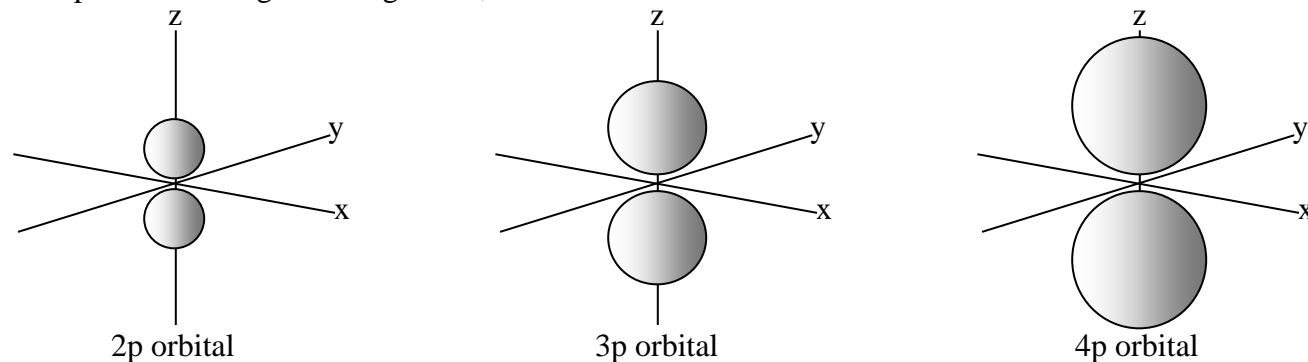
What is meant by a “probability location”? Recall that in the wave-mechanical model of the atom, electrons are not conceived of as “negatively charged points” in space that orbit the nucleus but rather as “negatively charged clouds” that surround the nucleus. The orbitals represent the region outside the nucleus where these clouds exist. Furthermore, each type of orbital (s, p, d, f) has its own unique shape and its own orientation in space with respect to the nucleus.

All s-orbitals have a spherical shape with the atomic nucleus at the center. 1s orbitals are the smallest, with the 2s orbitals being larger, and the 3s orbitals being even larger still, and etcetera.



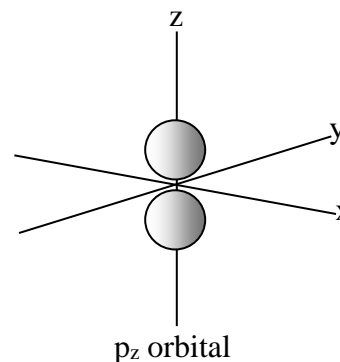
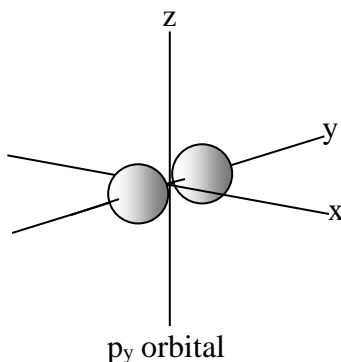
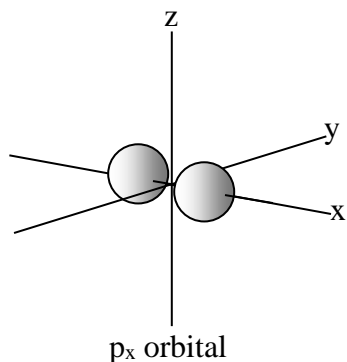
The region contained within the spherical shape of the s-orbital represents 90% probability of finding an s-electron there. The atomic nucleus is found at the “origin” of the three dimensional Cartesian axis.

All p-orbitals have a “dumbbell” shape. 2p orbitals are the smallest, with the 3p orbitals being larger, and the 4p orbitals being even larger still, and etcetera.

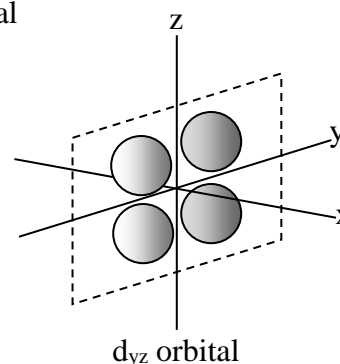
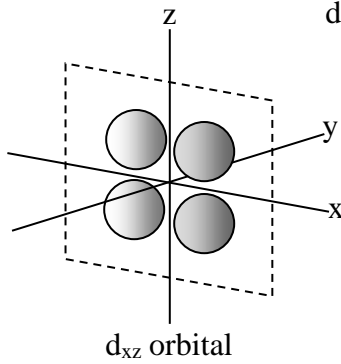
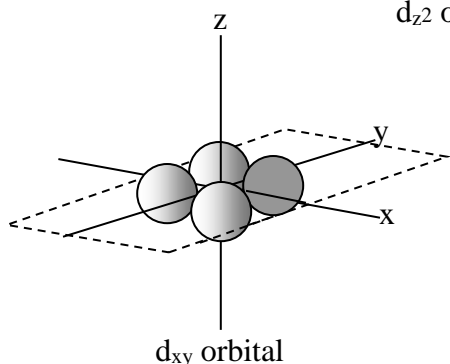
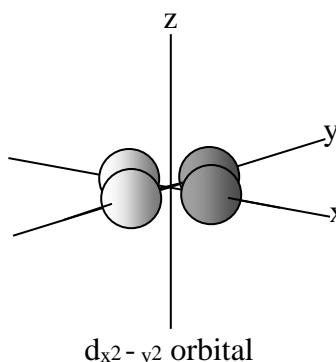
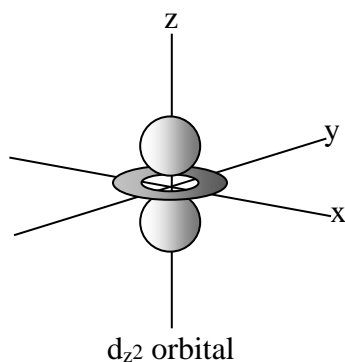


The region contained within the “dumbbell” shape of the p-orbital represents 90% probability of finding a p-electron there. The atomic nucleus is found at the “origin” of the three dimensional Cartesian axis.

Recall that each PEL (starting with the 2nd PEL) has three p-orbitals! Each of these three p-orbitals has a dumbbell shape **AND** each of these three p-orbitals has a unique orientation around the nucleus.



All five d-orbitals have their own unique orientation around the nucleus. Four of the d-orbitals have a “four-leafed clover” shape. The fifth d-orbital has a “dumbbell with a donut shape”. 3d orbitals are the smallest, with the 4d orbitals being larger, and the 5d orbitals being even larger still, and etcetera.



The region contained within the d-orbital shapes represents 90% probability of finding a d-electron there. The atomic nucleus is found at the “origin” of the three dimensional Cartesian axis.

The f-orbital shapes and orientations are too complex to represent here. 4f orbitals are the smallest, with the 5f orbitals being larger, and the 6f orbitals being even larger still, and etcetera. All seven f-orbital have their own unique orientation around the nucleus. The region contained within the f-orbital shapes represents 90% probability of finding an f-electron there

Every orbital, whether it is an s-orbital, a p-orbital, a d-orbital, or an f-orbital, can hold a maximum of two electrons. **Thus...**

Every **s-sublevel** is composed of 1 s-orbital which can hold 2 electrons and thus **2 electrons** total.

Every **p-sublevel** is composed of 3 p-orbitals which can hold 2 electrons each and thus **6 electrons** total.

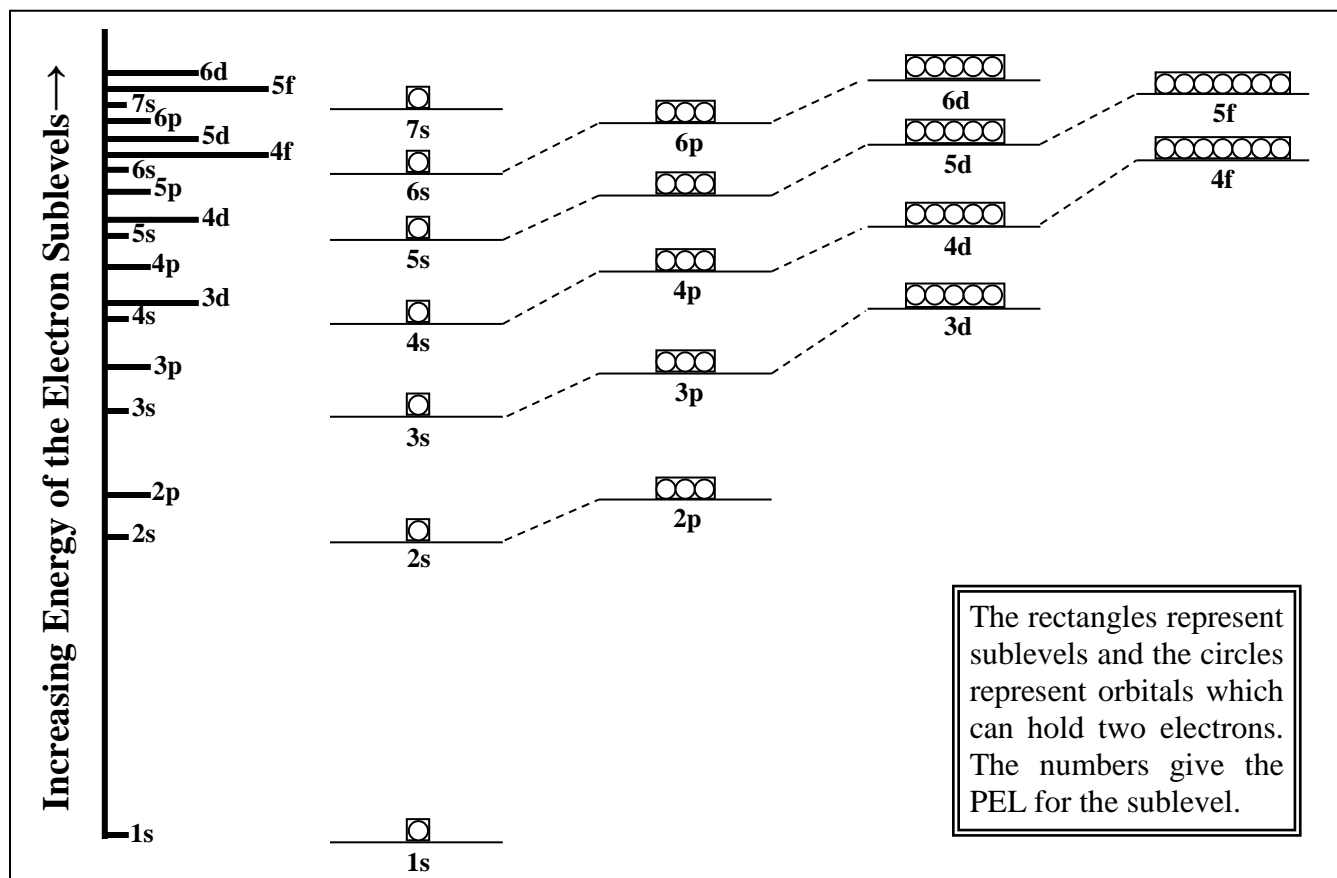
Every **d-sublevel** is composed of 5 d-orbitals which can hold 2 electrons each and thus **10 electrons** total.

Every **f-sublevel** is composed of 7 f-orbitals which can hold 2 electrons each and thus **14 electrons** total.

Finally, we can begin to answer the question,

Why is iron's shell electron configuration 2-8-14-2 and not 2-8-16?

It turns out that, in terms of increasing energy, the PEL's in atoms overlap as shown in the diagram below. Notice that the 4s orbital has less energy than the 3d orbital. Thus the 3rd PEL and the 4th PEL overlap. Electrons will begin to fill the 4th PEL before the 3rd PEL is full.



Notice that all orbitals in the same sublevel have the same level of energy. The first rule for building up an atom's electron configuration is the **Aufbau principle**: *an electron occupies the lowest-energy orbital that can receive it*. (Aufbau means 'building up' in German)

Since iron has 26 electrons let's use the diagram above to build up its electron configuration starting with the lowest-energy orbital... the 1s orbital.

There are a total of twenty six electrons in an atom of iron.

The single 1s orbital can hold 2 electrons and there are twenty four electrons left.

The single 2s orbital can hold 2 electrons and there are twenty two electrons left.

The three 2p orbitals can hold 6 electrons and there are sixteen electrons left.

The single 3s orbital can hold 2 electrons and there are fourteen electrons left.

The three 3p orbitals can hold 6 electrons and there are eight electrons left.

The single 4s orbital can hold 2 electrons and there are six electrons left.

The five 3d orbitals can hold ten electrons but there are only 6 electrons.

Now let's add up the electrons in each PEL:

The 1st PEL has a total of 2 electrons [2e⁻ in the 1s].

The 2nd PEL has a total of 8 electrons [2e⁻ in the 2s & 6e⁻ in the 2p].

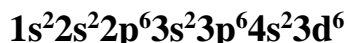
The 3rd PEL has a total of 14 electrons [2e⁻ in the 3s & 6e⁻ in the 3p & 6e⁻ in the 3d].

The 4th PEL has a total of 2 electrons [2e⁻ in the 4s].

Thus iron's **shell electron configuration** is: **2-8-14-2**

D) Electron Configurations:

Representing iron's shell electron configuration as **2-8-14-2** only indicates what PEL's are occupied but not what sublevels are occupied. There is another model for an iron atom's electron configuration:



[the superscripts represent the number of electrons in the orbital]

One reads this aloud by saying "*one ess two, two ess two, two pee six, three ess two, three pee six, four ess two, three dee six*".

From now on... if you are asked to write the electron configuration for an atom, this is the model to be used. Using the diagram on the previous page and the Aufbau principle, you can see that the electron configuration for ...

Hydrogen is: **1s¹** [1e⁻ in the 1s]

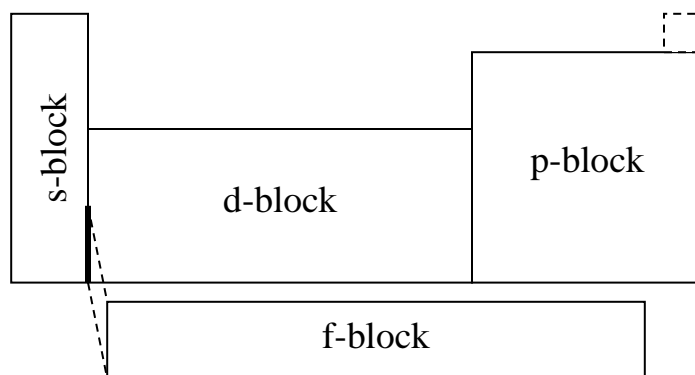
Helium is: **1s²** [2e⁻ in the 1s]

Lithium is: **1s²2s¹** [2e⁻ in the 1s & 1e⁻ in the 2s]

Sodium is: **1s²2s²2p⁶3s¹** [2e⁻ in the 1s & 2e⁻ in the 2s & 6e⁻ in the 2p & 1e⁻ in the 3s]

Now you may be wondering how you are to construct an atom's electron configuration without the use of the energy diagram on the previous page. Actually, it's fairly easy!!!

It turns out that the way that the Periodic Table is structured is not random. The Periodic Table is made up of four blocks as show below. Each block corresponds to the four lowest energy sublevels (*helium is part of the s-block*). The block in which an elements is found indicates the highest energy sublevel in an atom of that element. Iron is found in the d-block. You will confirm that iron's highest energy sublevel is the 3d-sublevel (**1s²2s²2p⁶3s²3p⁶4s²3d⁶**)!



The numbering of the rows in each block is shown below. Notice that the first row in the p-block is 2p because the first PEL that contains a p-sublevel is the 2nd PEL. The first row in the d-block is 3d because the first PEL that contains a d-sublevel is the 3rd PEL. The first row in the f-block is 4f because the first PEL that contains an f-sublevel is the 4th PEL.

1s		
2s		2p
3s		3p
4s	3d	4p
5s	4d	5p
6s	5d	6p
7s	6d	7p
	4f	
	5f	

Notice that the s-block is 2 elements wide and that it takes 2 electrons to fill an s-sublevel. Notice that the p-block is 6 elements wide and that it takes 6 electrons to fill a p-sublevel. Notice that the d-block is 10 elements wide and that it takes 10 electrons to fill a d-sublevel. Notice that the f-block is 14 elements wide and that it takes 14 electrons to fill an f-sublevel. This is not a coincidence!!! The elements may be placed on the Periodic Table in increasing order of their number of protons but the structure of the Periodic Table is base of the electronic structure of the elements!!!

Using the diagram below (which with practice I am sure you will be able to see in your head when looking at a normal Periodic Table) you can now construct the **electron configuration** for almost any element. This is how it's done...

[illegible]

Let's construct the **electron configuration** for lithium:

Step 1: Find lithium on the Periodic Table. The atomic number (number of protons) will be equal to the total number of electrons in the electron configuration. **3 electrons**

Step 2: Using the atomic numbers of the elements as a guide, move from #1 to #3 on the diagram above writing down the orbital designations as you go: 1s-1s-2s.

Step 3: Now write the electron configuration in its proper form: $1s^2 2s^1$
[Make sure that the superscripts sum to 3.]

Let's construct the **electron configuration** for chlorine:

Step 1: Find chlorine on the Periodic Table. The atomic number (number of protons) will be equal to the total number of electrons in the electron configuration. **17 electrons**

Step 2: Using the atomic numbers of the elements as a guide, move from #1 to # 17 on the diagram above writing down the orbital designations as you go:
1s-1s-2s-2s-2p-2p-2p-2p-2p-3s-3s-3p-3p-3p-3p-3p.

Step 3: Now write the **electron configuration** in its proper form: $1s^2 2s^2 2p^6 3s^2 3p^5$
[Make sure that the superscripts sum to 17.]

Let's construct the **electron configuration** for iron:

Step 1: Find iron on the Periodic Table. The atomic number (number of protons) will be equal to the total number of electrons in the electron configuration. **26 electrons**

Step 2: Using the atomic numbers of the elements as a guide, move from #1 to # 26 on the diagram above writing down the orbital designations as you go:
 1s-1s-2s-2s-2p-2p-2p-2p-2p-2p-3s-3s-3p-3p-3p-3p-3p-3p-4s-4s-3d-3d-3d-3d-3d-3d.

Step 3: Now write the electron configuration in its proper form: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$
[Make sure that the superscripts sum to 26.]

With practice, you can do step-2 without writing it down and move right to step-3.

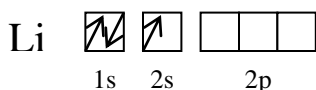
Iron's **electron configuration**: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ conveys a lot of information about the electrons in the atom. **First**: the electrons are found in four different PEL's. **Second**: the electrons are found in seven distinct sublevels. **Third**: the electrons with the highest energy are found in the 3d sublevel (the highest energy electrons will always appear on the end). **Fourth**: the third and fourth principal energy levels have overlapped since the 4s-sublevel is before the 3d-sublevel.

E) Electron Orbital Notation:

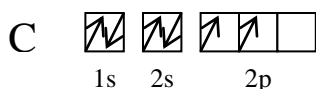
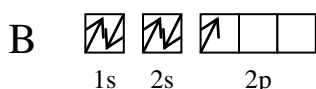
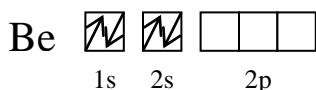
Iron's **electron configuration**: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ conveys a lot of information about the electrons in the atom but it does not give any information as to how the electrons are distributed in the orbitals. Thus a new type of electron configuration was developed which is called **electron orbital notation** or simply **orbital notation**. Recall, that an **orbital** is a probability location for an electron in an atom and that all s-sublevels contain one orbital, all p-sublevels contain three orbitals, d-sublevels contain five orbitals, and f-sublevels contain seven orbitals and that each and every orbital can hold up to two electrons.

To represent the orbitals, we will use a box. To represent the electrons, we will use arrows. Before we place the electrons (arrows) into the model, we need to learn two new rules. **Hund's rule** states that *orbitals of equal energy (in the same sublevel) must all be occupied with one electron before any one orbital is occupied by a second electron*. The **Pauli exclusion principle** states that *no two electrons in the same orbital can have the same spin*. Spin is a physical state of an electron in an orbital but which is not a property of the orbital itself. An electron can possess one of two possible spins which are opposite in their nature. To represent the two possible spins, we will use an "up arrow" \uparrow and a "down arrow" \downarrow . By convention, "up arrows" are entered first.

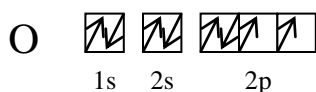
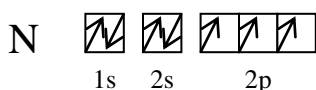
In order to better understand the model, the **orbital notations** for the second-period elements have been completed below.



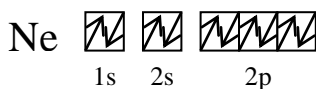
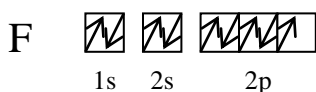
Notice that no two electrons in any one orbital have the same spin.
This is the **Pauli exclusion principle**.



Notice that the two electrons in the 2p sublevel of carbon occupy their own orbital.
This is **Hund's rule**.



Notice the pattern of how electrons are entered into the orbitals of the oxygen atom:
up, down, up, down, up, up, up, down



Using iron's electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$ we now can construct iron's orbitals notation as follows:



F) Condensed Electron Configurations and Orbital Notations:

As you can imagine, electron configurations and orbital notations for large elements can be quite long. Thus a condensed form for these models has been developed. The **condensed electron configuration** for iron is:



And the **condensed orbital notation** for iron is:



In the **condensed form**, the symbol for the noble gas from the previous period on the Periodic Table is used to replace the beginning of the model. Since iron is found in period-4 on the Periodic Table, then the noble gas from period-3, Ar, is used. Since the electron configuration for argon is $1s^2 2s^2 2p^6 3s^2 3p^6$, then this portion of iron's electron configuration can be replaced with [Ar]. Thus, iodine's electron configuration in the condensed form is: $[\text{Kr}]5s^2 4d^{10} 5p^5$.

G) Valence and Core Electrons:

In an atom, the outer most principal energy level is referred to as the **valence** and any electrons in the outer most principal energy level are called **valence electrons**. Typically, only electrons found in s and p-sublevels are valence electrons. Thus, the number of valence electrons in an atom never exceeds eight. Valence electrons are the ones an atom typically uses in bonding.

In an atom, the nucleus and any inner principal energy levels are referred to as the **core**. Any electrons found in the core are called **core electrons**.

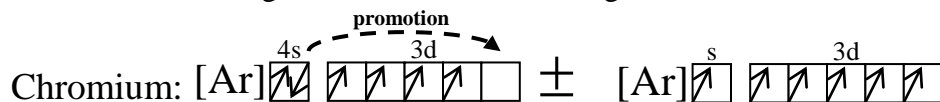
Using iron's electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^2} 3d^6$ we can see that there are 2 valence electrons (the 2 electrons in the 4s-sublevel) and that there are 24 core electrons. Notice that the valence electrons for iron are not found at the end of its electron configuration.

Using iodine's electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 \underline{5s^2} 4d^{10} \underline{5p^5}$ we can see that there are 7 valence electrons (the 2 electrons in the 5s-sublevel and the 5 electrons in the 5p-sublevel) and that there are 46 core electrons. Notice that valence electrons for iodine are in two different sublevels.

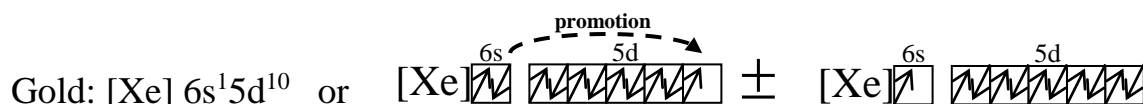
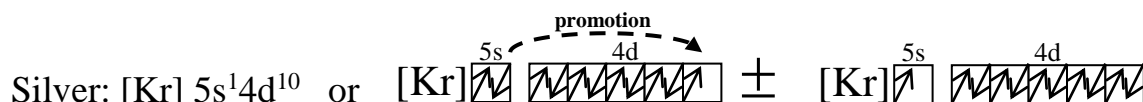
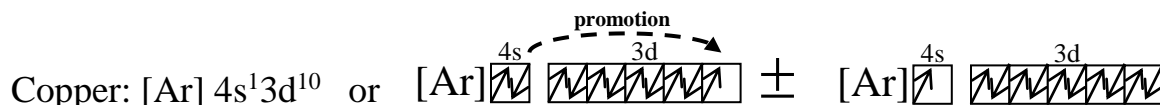
Sometimes, iron's electron configuration is written this way: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 \underline{4s^2}$. In this method, the electron configuration keeps all sublevels in the same PEL's together: $1s^2 / 2s^2 2p^6 / 3s^2 3p^6 3d^6 / \underline{4s^2}$ and thus will have all valence electrons at the end which does make it easier to find the valence electrons and the core electrons. However, this method ignores the ordering of the sublevels by increasing energy (*Aufbau principle*). It's also more difficult to use the Periodic Table to construct (*build up*) the electron configuration. For these reasons, I use the first method when writing electron configurations. I only mention this second method, because you may run across it and will then need to understand it.

H) Anomalous Electron Configurations and Orbital Notations:

Using your old Regents Periodic Table, you can find exceptions to the Aufbau principle in the d-block: Cr, Cu, Nb, Mo, Ru, Rh, Pd, Ag, Pt, and Au. In all of these cases, there should be 2 valence electrons if the Aufbau principle was being followed. **Remember, "mother nature" does not have to follow our rules!** For example chromium should be 2-8-12-2 or $1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^2} 3d^4$ (the valence has been underlined). However, chromium's actual electron configuration is: 2-8-13-1 or $1s^2 2s^2 2p^6 3s^2 3p^6 \underline{4s^1} 3d^5$.



In addition to chromium, you will be expected to know the exceptions for copper, silver, and gold.



Recall that **ions** are particles whose total number of electrons present is different from the total number of protons present. Recall, that a positive ion is formed when one or more electrons are lost and that they are taken out of the valence first.

or



$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$



or
 $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$



and

				
1s	2s	2p	3s	3p

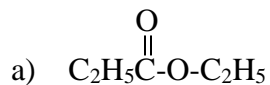
- 1) Using structural formulas, write a balanced equation for the esterification reaction between propanol and ethanoic acid. Name the ester produced.

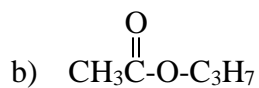
- 2) Using structural formulas, write a balanced equation for the esterification reaction between phenol and benzoic acid. Name the ester produced.

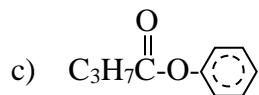
3) Using structural formulas, write a balanced equation for the esterification reaction between butanol and propanoic acid. Name the ester produced.

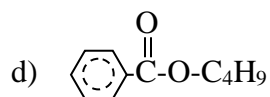
4) Using structural formulas, write a balanced equation for the esterification reaction between oxalic acid and two methanol molecules. Name the ester produced.

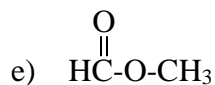
1) Name the following esters:

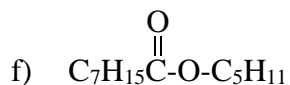












2) Write the structural formulas for the following esters:

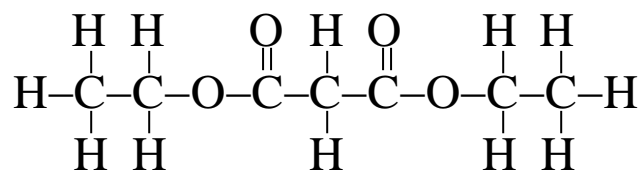
a) hexyl methanoate

b) phenyl ethanoate

c) ethyl propanoate

d) methyl benzoate

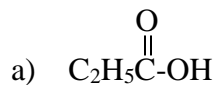
- 3) A diester is an ester of a dicarboxylic acid molecule and two alcohol molecules. If the diester shown below is called diethyl 1,3-propanedioate, then draw the structural formula for diethyl 1,5-pentanedioate.

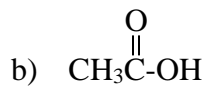


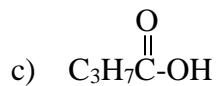
- 4) 1.0000 g of an unknown ester burns to give 2.2035 g of CO_2 and 0.9306 g H_2O . Determine the empirical formula of the unknown ester.

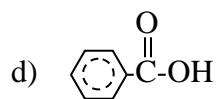
- 5) If the unknown ester is known to contain two oxygen atoms per molecule, the what is the molecular formula of the unknown ester?

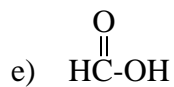
1) Name the following carboxylic acids:

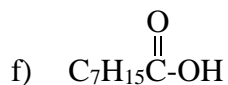












2) Write the structural formulas for the following carboxylic acids:

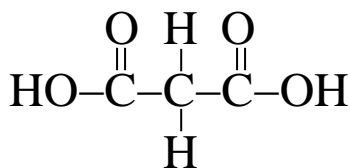
a) pentanoic acid

b) oxalic acid

c) acetic acid

d) hexanoic acid

- 3) A dicarboxylic acid is an acid molecule containing two carboxyl groups. If the dicarboxylic acid shown below is called propane-1,3-dioic acid, then draw the structural formula for dicarboxylic acid pentane-1,5-dioic acid.



- 4) As you recall, an acid's pK_a is an indication as to its strength. The monocarboxylic acid, ethanoic acid (CH_3COOH), has a pK_a of 4.74 and the dicarboxylic acid, ethanedioic acid (HOOCCOOH), has a first pK_a of 1.23 and a second pK_a of 4.19. Even the second pK_a of ethanedioic acid shows that its second carboxyl group is more acidic than ethanoic acid's only carboxyl group. In terms of their molecular structures, explain the difference in their acid strengths.

- 5) What are the common names for ethanoic acid and ethanedioic acid?

ethanoic acid: _____ & ethanedioic acid: _____

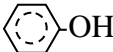
- 6) Citric acid, the acid found in orange juice, lemon juice, and grapefruit juice is a tricarboxylic acid. Its IUPAC name is 3-carboxy, 3-hydroxy, pentane-1,5-dioic acid. Draw the structural formula for citric acid.

1) Name the following alcohols:

a) $\text{CH}_3\text{CH}_2\text{OH}$ _____

b) CH_3OH _____

c) $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ _____

d)  _____

e) $\text{CH}_3\text{CH}(\text{OH})\text{CH}_3$ _____

f) $\text{CH}_2(\text{OH})\text{CH}_2\text{OH}$ _____

2) Write the structural formulas for the following alcohols:

a) 1-butanol

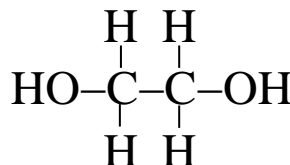
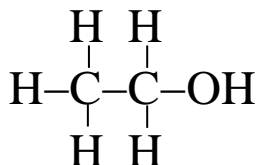
b) 2-butanol

c) glycerol

d) 3-pentanol

- 3) Even though alcohols contain the “–OH” group, they are not hydroxides. Explain the behavioral difference between an alcohol and a hydroxide in water.

- 4) Which of the two alcohols shown below would you expect to have the higher boiling point and why?



- 5) Gasoline is a mixture of different hydrocarbons one of the most important of which is 2,2,4-trimethyl pentane a form of octane (C_8H_{18}). Write a balanced equation, using whole number coefficients, for the complete combustion of 2,2,4-trimethyl pentane.

- 6) If the heat of combustion for the reaction you wrote in question-5 is “–10943 kJ” then calculate the heat of formation of $\text{C}_8\text{H}_{18(\text{l})}$ using the heats of formation of the other substances. Assume that the water is $\text{H}_2\text{O}_{(\text{l})}$

- 1) What percent of a simple cubic unit cell is occupied by atoms? _____
- 2) What percent of a body-centered cubic unit cell is occupied by atoms? _____
- 3) What percent of a face-centered cubic unit cell is occupied by atoms? _____
- 4) You are given a small bar of an unknown metal X. You find that metal X has a density of 10.5 g/cm³. An X-ray diffraction crystallography experiment determines that metal X has a face-centered cubic unit cell structure whose edge length is 4.09 Å (1 Å = 1×10⁻¹⁰ m). Using molar mass, identify metal X.

Complete the following questions from the textbook:

11.58

11.60

a) _____ (*empirical formula*)

11.62

a) _____ b) _____

c)

d)

11.93

a)

b)

Lattice energy is the energy required to completely separate (decompose) 1 mole of an ionic compound from its crystal lattice (solid state) into free gaseous ions. An ionic compound's lattice energy directly relates to the strength of the ionic bonding present.
The greater the value of the lattice energy, the stronger the ionic bond is.

Lattice energy cannot be directly determined through experiment. It can, however, be determined from a series of well defined steps known as the Born-Haber cycle.

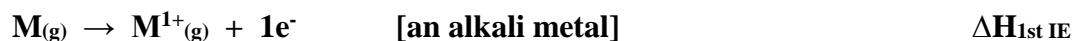
The Born-Haber cycle is a thermochemical cycle, based on Hess's law, which relates an ionic compound's lattice energy to its heat of formation.

Use the following steps and the data on the following pages to determine the lattice energy (ΔH_{LE}) for an alkali metal-halogen compound (MX) or for an alkaline earth metal-halogen compound (MX_2) given the heat of formation (ΔH_f):

Step-1: Vaporize the Solid Metal to Gaseous Metallic Atoms



Step-2: Ionize the Gaseous Metallic Atoms to Gaseous Metallic Ions



or



Step-3: Decompose the Halogen Molecules to Gaseous Halogen Atoms



or



Step-4: Ionize the Gaseous Halogen Atoms to Gaseous Halogen Ions



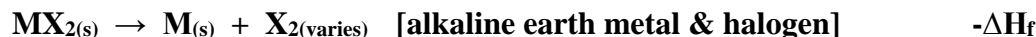
or



Step-5: Reverse the Given Formation Reaction and Negate the Heat of Formation



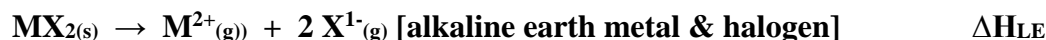
or



Step-6: To Determine the Lattice Energy (ΔH_{LE}), Add Steps 1-5



or



(Use a pencil... just in case...)

This image shows a single sheet of white paper with horizontal ruling lines. The lines are evenly spaced and run across the width of the page. There are no margins, text, or other markings on the paper.

Use this data to complete HW EP.1

Substance or Ion	ΔH_f° (kJ/mol)	S° (J/mol·K)	Substance or Ion	ΔH_f° (kJ/mol)	S° (J/mol·K)	Substance or Ion	ΔH_f° (kJ/mol)	S° (J/mol·K)
Ag(s)	0.0	42.7	ClO ₃ ¹⁻ (aq)	-98.3	162.3	H ₂ (g)	0.0	130.6
Ag ¹⁺ (aq)	+105.9	73.9	ClO ₄ ¹⁻ (aq)	-131.4	182.0	I ₂ (s)	0.0	116.7
AgBr(s)	-99.5	107.1	Cr(s)	0.0	23.8	I(g)	106.6	180.7
AgCl(s)	-127.0	96.1	CrO ₄ ²⁻ (aq)	-863.2	38.5	I ¹⁻ (aq)	-55.9	109.4
AgI(s)	-62.4	114.2	Cr ₂ O ₃ (s)	-1128.4	81.2	K(s)	0.0	63.6
Ag ₂ O(s)	-30.6	121.7	Cs(s)	0.0	85.15	K(g)	90.9	160.2
Ag ₂ S(s)	-31.8	145.6	Cs(g)	76.7	175.5	K ¹⁺ (aq)	-251.2	102.5
Al(s)	0.0	28.3	Cs ¹⁺ (aq)	-248.0	133.0	KBr(s)	-392.2	96.4
Al ³⁺ (aq)	-524.7	-313.4	CsCl(s)	-442.8	101.2	KCl(s)	-435.9	82.7
Al ₂ O ₃ (s)	-1669.8	51.0	CsBr(s)	-395.0	121.0	KClO ₃ (s)	-391.4	143.0
Ba(s)	0.0	67	CsF(s)	-554.7	88.0	KF(s)	-562.6	66.6
Ba(g)	175.6	144.8	Cu(s)	0.0	33.3	KI(s)	-328	106.4
Ba ²⁺ (aq)	-538.4	13	Cu(g)	338.4	166.3	Li(s)	0.0	29.1
BaCl ₂ (s)	-860.1	126	Cu ²⁺ (aq)	+64.4	-98.7	Li(g)	161	138.7
BaCO ₃ (s)	-1218.8	112.1	CuO(s)	-155.2	43.5	Li ¹⁺ (aq)	-278.5	14
BaO(s)	-558.1	70.3	Cu ₂ O(s)	-166.7	100.8	LiF(s)	-616.9	14
BaSO ₄ (s)	-1465.2	132.2	CuS(s)	-48.5	66.5	LiCl(s)	-408	59.3
Br ₂ (l)	0.0	152.3	CuSO ₄ (s)	-769.9	113.4	LiBr(s)	-351	74.1
Br(g)	111.8	174.9	F ₂ (g)	0.0	203.3	LiI(s)	-270	85.8
Br ¹⁻ (aq)	-120.9	80.7	F(g)	80.0	158.7	Mg(s)	0.0	32.5
C(s) graphite	0.0	5.7	F ¹⁻ (aq)	-329.1	-9.6	Mg(g)	147.2	148.6
C(g)	718.4	672.9	Fe(s)	0.0	27.2	Mg ²⁺ (aq)	-462.0	-118.0
CCl ₄ (l)	-139.5	214.4	Fe(g)	415.5	180.5	MgCl ₂ (s)	-641.8	89.5
CH ₄ (g)	-74.8	186.2	Fe ²⁺ (aq)	-87.9	-113.4	MgCO ₃ (s)	-1113	65.7
CHCl ₃ (l)	-131.8	202.9	Fe ³⁺ (aq)	-47.7	-293.3	MgO(s)	-601.8	26.8
CH ₃ OH(l)	-238.6	126.8	Fe ₂ O ₃ (s)	-822.2	90.0	Mg(OH) ₂ (s)	-924.7	63.1
CO(g)	-100.5	197.9	Fe ₃ O ₄ (s)	-1120.9	146.4	MgSO ₄ (s)	-1278.2	91.6
CO ₂ (g)	-393.5	213.6	H ₂ (g)	0.0	130.6	Mn(s)	0.0	203.3
CO ₃ ²⁻ (aq)	-676.3	-53.1	H(g)	217.9	114.6	Mn(g)	280.7	173.6
C ₂ H ₂ (g)	+226.7	200.8	H ¹⁺ (aq)	0.0	0.0	Mn ²⁺ (aq)	-462.0	-73.6
C ₂ H ₄ (g)	+52.3	219.5	HBr(g)	-36.2	198.5	MnO(s)	-384.9	60.2
C ₂ H ₆ (g)	-84.7	229.5	HCO ₃ ¹⁻ (aq)	-691.1	95.0	MnO ₂ (s)	-519.7	53.1
C ₃ H ₈ (g)	-103.8	269.9	HCl(g)	-92.3	186.7	MnO ₄ ¹⁻ (aq)	-518.4	190.0
n-C ₄ H ₁₀ (g)	-124.7	310.0	HF(g)	-268.6	173.5	Na(s)	0.0	51.0
n-C ₅ H ₁₂ (l)	-173.1	262.8	HI(g)	+25.9	206.3	Na(g)	107.7	153.7
C ₂ H ₅ OH(l)	-277.6	160.7	HNO ₃ (l)	-173.3	155.6	Na ¹⁺ (aq)	-239.7	60.2
Ca(s)	0.0	41.6	H ₂ O(g)	-241.8	188.7	NaCl(s)	-411.0	72.4
Ca(g)	179.3	154.8	H ₂ O(l)	-285.8	69.9	NaF(s)	-569.0	58.6
Ca ²⁺ (aq)	-543.0	-55.2	H ₂ O ₂ (l)	-187.6	88.6	NaOH(s)	-426.7	52.3
CaCl ₂ (s)	-795.0	113.8	H ₃ PO ₄ (aq)	-1288.3	158.2	N ₂ (g)	0.0	191.5
CaCO ₃ (s)	-1207.0	92.9	H ₂ PO ₄ ¹⁻ (aq)	-1302.5	89.1	N(g)	472.7	153.3
CaF ₂ (s)	-1215.0	68.9	HPO ₄ ²⁻ (aq)	-1298.7	-36.0	NH ₃ (g)	-46.2	192.5
CaO(s)	-635.5	39.7	H ₂ S(g)	-20.1	205.6	NH ₄ ¹⁺ (aq)	-132.8	112.8
Ca(OH) ₂ (s)	-986.6	76.1	H ₂ SO ₄ (l)	-811.3	156.9	NH ₄ Cl(s)	-315.4	94.6
CaSO ₄ (s)	-1432.7	106.7	HSO ₄ ¹⁻ (aq)	-885.8	126.9	NH ₄ NO ₃ (s)	-365.1	151.0
Cl ₂ (g)	0.0	222.9	Hg(l)	0.0	77.4	NO(g)	+90.4	210.6
Cl(g)	121.0	158.0	HgO(s)	-90.7	72.0	NO ₂ (g)	+33.9	240.5
Cl ¹⁻ (aq)	-167.4	55.1	HgS(s)	-58.2	77.8	NO ₃ ¹⁻ (aq)	-206.6	146.4

Use this data to complete HW EP.1

Substance or Ion	ΔH_f° (kJ/mol)	S° (J/mol·K)	Substance or Ion	ΔH_f° (kJ/mol)	S° (J/mol·K)	Substance or Ion	ΔH_f° (kJ/mol)	S° (J/mol·K)
Ni(s)	0.0	30.1	PCl _{3(g)}	-306.4	311.7	Sn(s)	0.0	51.5
Ni ²⁺ (aq)	-64.0	-159.4	PCl _{5(g)}	-398.9	352.7	Sn ²⁺ (aq)	-10.0	-24.7
NiO(s)	-244.3	38.6	PO ₄ ³⁻ (aq)	-1284.1	-218	SnCl _{2(s)}	-349.8	122.6
O _{2(g)}	0.0	205.0	Rb(s)	0.0	69.5	SnCl _{4(l)}	-545.2	258.6
O(g)	247.5	161.0	Rb(g)	85.8	170.0	SnO(s)	-286.2	56.5
OH ¹⁻ (aq)	-229.9	-10.5	RbI(s)	-328.0	118.0	SnO _{2(s)}	-580.7	52.3
Pb(s)	0.0	64.9	S(s)	0.0	31.9	Sr(s)	0.0	54.4
Pb ²⁺ (aq)	+1.6	21.3	S ²⁻ (aq)	+41.8	22.2	Sr(g)	164.0	---
PbBr _{2(s)}	-277.0	161.5	SO _{2(g)}	-296.9	248.5	Sr ²⁺ (aq)	-545.5	-39.0
PbCl _{2(s)}	-359.2	136.4	SO _{3(g)}	-395.2	256.2	Zn(s)	0.0	41.6
PbO(s)	-217.9	69.5	SO ₄ ²⁻ (aq)	-907.5	17.2	Zn(g)	130.7	160.9
PbO _{2(s)}	-276.6	76.6	Si(s)	0.0	18.7	Zn ²⁺ (aq)	-152.4	-106.5
Pb ₃ O _{4(s)}	-734.7	211.3	Si(g)	368.2	167.8	ZnCl _{2(s)}	-415.1	111.5
P _{4(s)} white	0.0	41.1	SiCl _{4(l)}	-640.1	239.3	ZnO(s)	-348.0	43.9
P(g)	316.4	163.2	SiO _{2(s)}	-859.4	41.8	ZnS(s)	-202.9	57.5

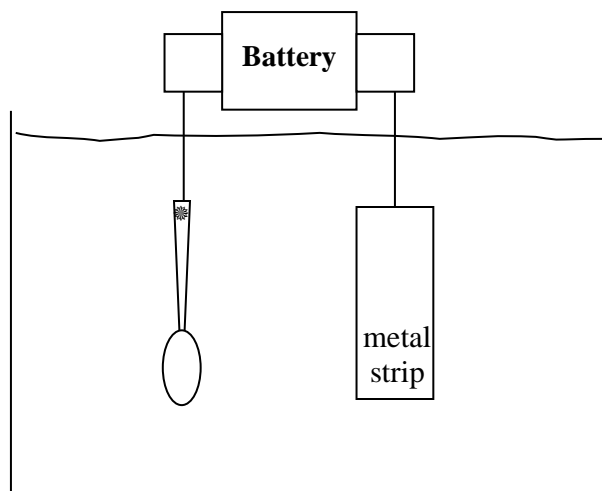
1312	1 st Ionization Energy [1 st IE] (kJ/mol)															2372	
520	900											801	1086	1402	1314	1681	2081
496	736											578	787	1012	1000	1251	1521
419	590	633	659	651	653	717	762	760	737	745	906	579	762	944	941	1140	1351
403	549											558	708	831	869	1008	1170
376	503											589	715	703	812	—	1037
393	—																
<u>2nd IE values for alkaline earth metals in kJ/mol:</u>																	
Be = 1760, Mg = 1450, Ca = 1145, Sr = 1064, Ba = 965																	

[illegible][illegible]

Lattice Energies for Some Ionic Compounds					
Compound	(kJ/mol)	Compound	(kJ/mol)	Compound	(kJ/mol)
LiF	1030	NaI	682	MgCl ₂	2326
LiCl	834	KF	808	SrCl ₂	2127
LiI	730	KCl	701	MgO	3795
NaF	910	KBr	671	CaO	3414
NaCl	788	CsCl	657	SrO	3217
NaBr	732	CsI	600	ScN	7547

Single Bond	Bond Enthalpy	Double Bond	Bond Enthalpy	Triple Bond	Bond Enthalpy
C-C	347 kJ/mol	C=C	612 kJ/mol	C≡C	820 kJ/mol
C-O	351 kJ/mol	C=O	745 kJ/mol	C≡O	1075 kJ/mol
C-H	414 kJ/mol	C=O in CO ₂	799 kJ/mol	C≡N	890 kJ/mol
N-H	389 kJ/mol	N=O	607 kJ/mol	N≡N	941 kJ/mol
S-O	265 kJ/mol	O=O	495 kJ/mol		
O-O	146 kJ/mol	S=O	498 kJ/mol		
O-H	464 kJ/mol	N=O	607 kJ/mol		
H-H	436 kJ/mol				

- 1) The following figure represents an electroplating device that will plate copper onto a stainless steel spoon.
- Label the positive pole of the battery with a “+”.
 - Label the negative pole of the battery with a “-”.
 - Label the metal strip with an appropriate element symbol.
 - Label the cathode.
 - Label the anode.
 - Label the solution with an appropriate compound for the electrolyte.



- 2) (i) Predict the products of the electrolysis of 1 M $\text{ZnBr}_{2(aq)}$ and (ii) write the appropriate half-reactions.

at the cathode: (i) _____ (ii) _____

at the anode: (i) _____ (ii) _____

Write the equation for the net ionic reaction of the electrolysis of 1 M $\text{ZnBr}_{2(aq)}$.

Net Ionic Reaction: _____

- 3) (i) Predict the products of the electrolysis of 1 M $\text{Ba(OH)}_{2(aq)}$ and (ii) write the appropriate half-reactions.

at the cathode: (i) _____ (ii) _____

at the anode: (i) _____ (ii) _____

Write the equation for the net ionic reaction of the electrolysis of 1 M $\text{Ba(OH)}_{2(aq)}$.

Net Ionic Reaction: _____

- 4) (i) Predict the products of the electrolysis of 1 M $\text{NaI}_{(\text{aq})}$ and (ii) write the appropriate half-reactions.

at the cathode: (i) _____ (ii) _____

at the anode: (i) _____ (ii) _____

Write the equation for the net ionic reaction of the electrolysis of 1 M $\text{NaI}_{(\text{aq})}$.

Net Ionic Reaction: _____

- 5) (i) Predict the products of the electrolysis of 1 M $\text{HNO}_{3(\text{aq})}$ and (ii) write the appropriate half-reactions.

at the cathode: (i) _____ (ii) _____

at the anode: (i) _____ (ii) _____

Write the equation for the net ionic reaction of the electrolysis of 1 M $\text{HNO}_{3(\text{aq})}$.

Net Ionic Reaction: _____

- 6) (i) Predict the products of the electrolysis of 1 M $\text{KF}_{(\text{aq})}$ and (ii) write the appropriate half-reactions.

at the cathode: (i) _____ (ii) _____

at the anode: (i) _____ (ii) _____

Write the equation for the net ionic reaction of the electrolysis of 1 M $\text{KF}_{(\text{aq})}$.

Net Ionic Reaction: _____

- 7) (i) Predict the products of the electrolysis of 1 M $\text{AlCl}_{3(\text{aq})}$ and (ii) write the appropriate half-reactions.

at the cathode: (i) _____ (ii) _____

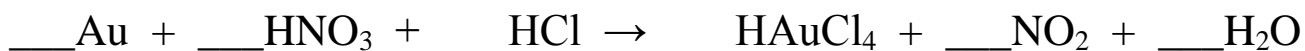
at the anode: (i) _____ (ii) _____

Write the equation for the net ionic reaction of the electrolysis of 1 M $\text{AlCl}_{3(\text{aq})}$.

Net Ionic Reaction: _____

Balance the following redox reactions:

- 7) *Aqua regia*, a mixture of 1-part concentrated nitric acid and 3-parts concentrated hydrochloric acid, was developed by alchemists as a means to “**dissolve**” gold [*gold will neither “dissolve” in nitric acid or hydrochloric acid alone*]. The process is actually a redox reaction with the unbalanced reaction shown below [the AuCl_4^{1-} in the HAuCl_4 is a complex ion]. **Balance the reaction.**



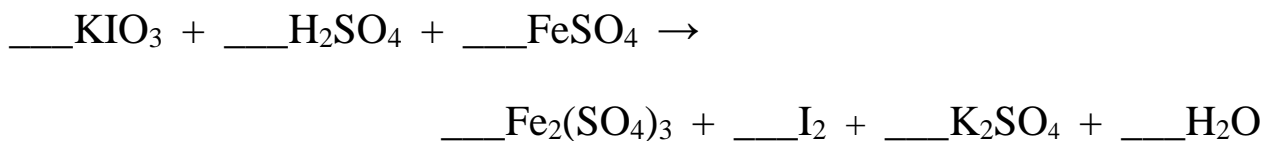
- 8) Unlike gold, mercury can be brought into solution by nitric acid alone. The process is shown by the following unbalanced redox reaction. Hg_2^{2+} is the polyatomic ion mercury (I). **Balance the reaction.**



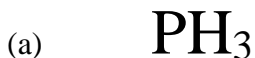
- 9) A 12 volt lead-acid-automobile battery consists of six voltaic cells in series, each producing 2 volts. The unbalanced reaction that takes place in each voltaic cell is shown below. **Balance the reaction.**



- 10) FeSO_4 is the most common ingredient used to make iron supplements. One method to test for the amount of iron found in a certain brand of iron supplement is to conduct a redox titration. The unbalanced reaction that takes place in the titration is found below. **Balance the reaction.**



1) Assign the oxidation number above each element in the following formulas.



2) Most metals react with most acids to produce a salt and hydrogen gas. Aluminum metal vigorously reacts with 1 M hydrochloric acid.

(a) Write the balanced equation for the reaction between aluminum metal and hydrochloric acid.

(b) What species is being oxidized? _____

(c) What species is being reduced? _____

3) Copper metal does not react with most acids. However, copper metal does react with nitric acid to produce colorless nitrogen monoxide gas. Immediately, the nitrogen monoxide gas reacts with oxygen gas in the air to produce brown nitrogen dioxide gas.

(a) Write the balanced equation for the reaction between nitrogen monoxide gas and oxygen gas.

(b) What is the oxidation number for the nitrogen in nitrogen monoxide gas? _____

(c) What is the oxidation number for the nitrogen in nitrogen dioxide gas? _____

4) Chromium can exist as the following species: Cr^0 , Cr^{+2} , Cr^{+3} , and Cr^{+6} .

(a) Which of these species can act as reducing agents? _____

(b) Which of these species can act as oxidizing agents? _____

(c) Write and name the formulas for two different **polyatomic ions** that contain the species Cr^{+6} .

_____, _____ ion & _____, _____ ion
FORMULA NAME FORMULA NAME

5) In terms of electrons, define oxidation.

6) When an element is oxidized, what happens to its oxidation number?

7) In terms of electrons, define oxidizing agent.

Complete the following questions from the textbook:

20.10) (a) _____

(b) _____

(c) _____

(d) _____

20.12) (a) _____

(b) _____

(c) _____

20.14) (a) _____

(b) _____

(c) _____

(d) _____

- 1) A 30.0 mL sample of 0.150 M $\text{HC}_2\text{H}_3\text{O}_2$ is titrated with 0.150 M NaOH. Calculate the pH after the following volumes of NaOH have been added and list the major species present at each pH:

(a) at 0.0 mL the pH = _____; the major species present are: _____

(b) at 15.0 mL the pH = _____; the major species present are: _____

(c) at 29.0 mL the pH = _____; the major species present are: _____

(d) at 30.0 mL the pH = _____; the major species present are: _____

(e) at 31.0 mL the pH = _____; the major species present are: _____

Complete the following questions from the textbook:

17.4)

(a) _____

(b) _____

(c) _____

(d) _____

17.32)

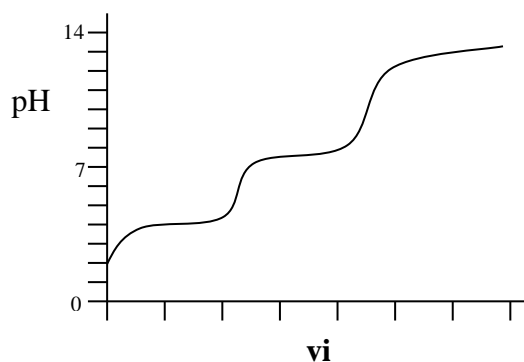
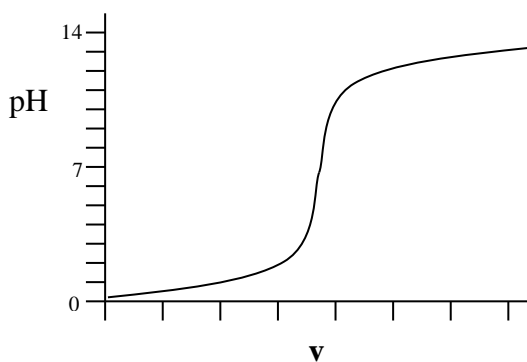
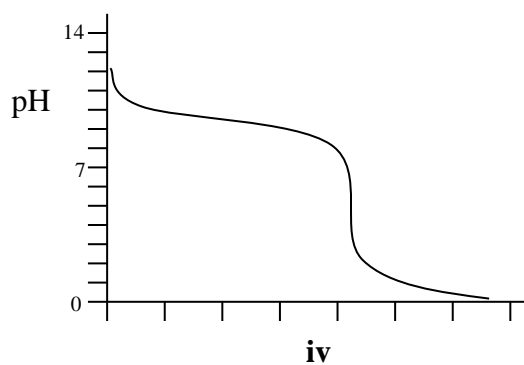
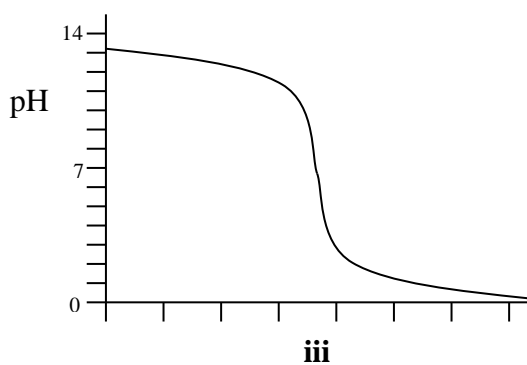
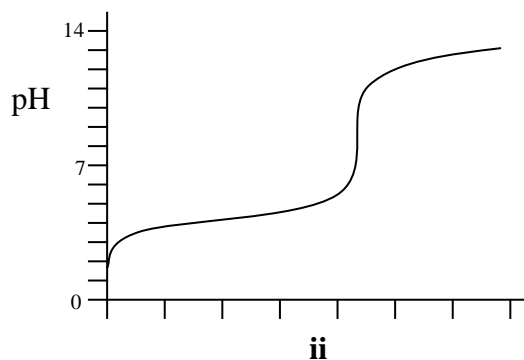
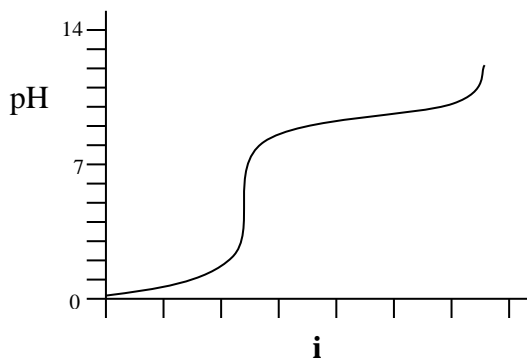
(a) _____

(b) _____

(c) _____

(d) _____

1) Match the type of titration to the sketch of its titration curve.



- a) Adding a strong acid to a strong base. _____
- b) Adding a strong base to a weak monoprotic acid. _____
- c) Adding a strong base to a polyprotic acid. _____
- d) Adding a strong acid to a weak base. _____

Complete the following questions from the textbook:

17.34)

a) _____

b) _____

c) _____

17.40)

a)

b)

c)

d)

e)

17.44)

a) _____

- 1) What is the molar concentration of Cl^- ions in a 5.00 mL sample of seawater if it takes 32.5 mL of 0.100M AgNO_3 to bring about the complete precipitation of the Cl^- ions?

- 2) Based on your answer to question-#1, how many grams of Cl^- ions would be present in one liter of the seawater?

- 3) How many milliliters would be needed to titrate 10.0 mL of 0.050M NaOH with 0.100 M $\text{Cu}(\text{NO}_3)_2$ (assume complete precipitation)?

- 4) What indicator would best signal the equivalence point during the titration of HCl and NaOH?

- 5) What indicator would best signal the equivalence point during the titration of HCl and NH_3 ?

- 6) What indicator would best signal the equivalence point during the titration of HF and NaOH?

- 7) Predict the pH of a solution if a sample turns methyl orange to orange and litmus to red.

Complete the following questions from the textbook:

17.38a)

17.38c)

17.64) K_{sp} values are found on page 1127 in the textbook.

K_{sp} values are found on page 1127 and K_f values are found on page 750 in the textbook

- 1) Which complex ion found in Table 17.1 in the textbook is the most stable?

- 2) Write the equation for the dissolution of solid $\text{Fe}(\text{OH})_2$ in an aqueous solution containing CN^{1-} ions.

- 3) Calculate the K_{eq} for the reaction of dissolving of solid $\text{Fe}(\text{OH})_2$ in an aqueous solution of NaCN.

- 4) Calculate the molar solubility of solid $\text{Fe}(\text{OH})_2$ in an aqueous solution of 1.00M NaCN.

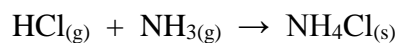
- 5) Write the equation for the dissolution of solid NiC_2O_4 in an aqueous solution containing NH_3 molecules. *Circle the major species that exist at equilibrium.*

Complete the following questions from the textbook:

17.58)

17.60)

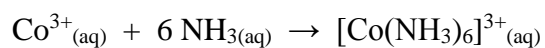
- 1) Explain how the following reaction is an acid-base reaction. Identify the acid and the base.



- 2) What is a coordinate covalent bond?

- 3) What is a complex ion?

- 4) Explain how the following reaction is an acid-base reaction. Identify the acid and the base.



- 5) State the charges for the following complex ions.

a) $[\text{Ag}(\text{CN})_2]^{??}$ _____

b) $[\text{Zn}(\text{H}_2\text{O})_6]^{??}$ _____

c) $[\text{MoCl}_6]^{??}$ _____ if Mo is a 3+ cation

d) $[\text{Ni}(\text{NH}_3)_6]^{??}$ _____ if Ni is a 2+ cation

- 6) Hydrated metallic cations are weak acids. Predict the products for the ionization of $[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ in water and write the equilibrium equation. *Circle the major species that exist at equilibrium.*

- 7) Hydrated metallic cations are weak acids. Predict the products for the ionization of $[\text{Cu}(\text{H}_2\text{O})_4]^{2+}$ in water and write the equilibrium equation. *Circle the major species that exist at equilibrium.*

Complete the following questions from the textbook:

16.104

f) _____

b) _____

c) _____

16.109

a) _____

b) _____

c) _____

- 1) Tin and lead form the alloy known as solder. Is the alloy substitutional or interstitial? Justify your answer. *You may need to research the atomic radii for tin and lead.*

- 2) Is a semiconductor of germanium doped with arsenic a p-type or an n-type semiconductor? Explain.

- 3) Explain how could you test experimentally if TiO_2 is an ionic solid or a network solid?

- 4) The band gap between the valance band and the conduction band in amorphous selenium, which is used in the xerographic photocopying process, is $2.9 \times 10^{-19} \text{ J}$?

- a) What is the longest wavelength (**in nm**) that can make amorphous selenium conduct electricity?
($E = hc/\lambda$)

- b) Can amorphous selenium be used in infrared burglar alarms? Explain.

- 5) DNA and RNA are a large polymeric molecules that are composed of two strands that wind around each other in double helix? Based on the structures of the nucleotide bases found on page 1103, explain how the two strands in the DNA and RNA helices are held to each other.

- 6) Describe the difference between quartz and quartz glass.

Complete the following questions from the textbook:

23.4

a)

23.24

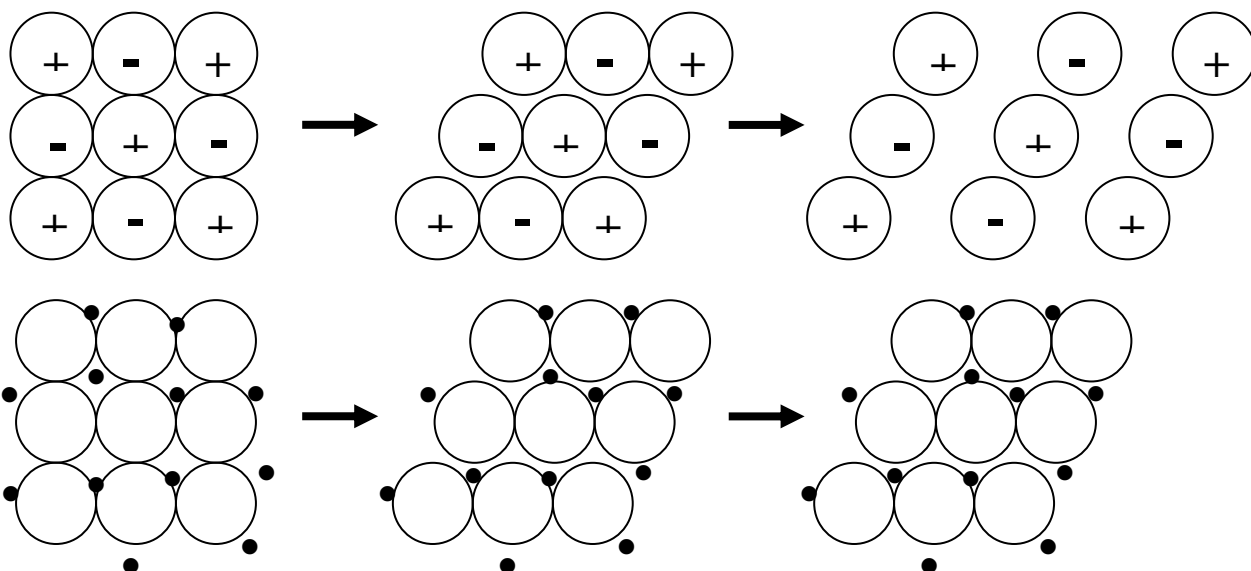
23.30

First question

Second question

23.62

- 1) Ionic solids are brittle whereas metallic solids are malleable. The diagrams below show an ionic solid and a metallic solid before and after being struck by a hammer. Explain, based on the diagrams and in terms of the bonding present, why ionic solids are brittle whereas metallic solids are malleable.



- 2) Dry ice is solid carbon dioxide. In a solid sample of dry ice...
- state the intramolecular forces present. _____
 - state the intermolecular forces present. _____
- 3) Explain why water vapor, at 220 atm of pressure, cannot be liquefied. (*Consult the phase diagram in the notes.*)
- _____
- _____
- 4) Explain why an ice skater is actually gliding on a thin layer of liquid water that forms on top of the ice just as the skater passes over it.
- _____
- _____
- _____
- _____
- _____

Complete the following questions from the textbook:

11.6

a) b.p. = _____ m.p. = _____

b) (i) _____ (ii) _____ (iii) _____

11.72

a) _____

b) _____

c) _____

d) _____

e) _____

f) _____

11.74

a) _____

b) _____

c) _____

d) _____

e) _____

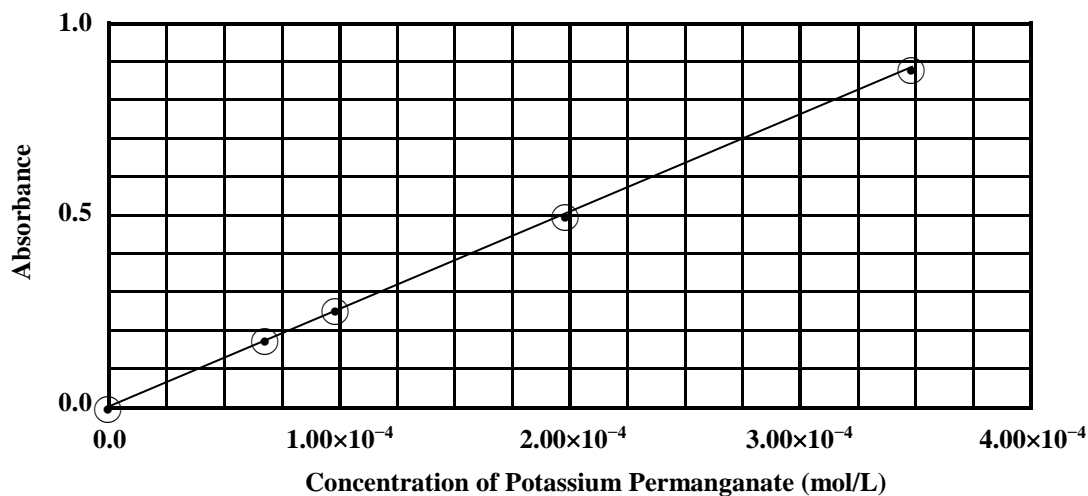
11.100

a)

b) _____ c) _____

d) _____

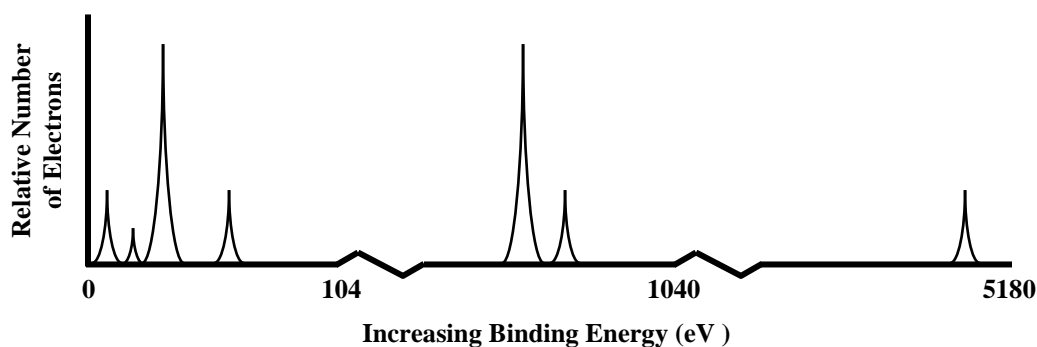
- 1) Based on the standard absorption curve below, what is the concentration of an unknown solution of potassium permanganate if it has an absorbance reading of 0.7 at 520 nm?



_____ mol/L [1]

- 2) If the path length of the solution through which the light passed was 1.0 cm, then calculate the value of the molar absorptivity coefficient of the permanganate ion at 520 nm. (*include units*) [2]

- 3) Identify the element whose photoelectron spectrum (PES) is shown below.



_____ [1]

- 4) The ionization energy for the single electron in hydrogen is 1312 kJ/mol. (a) Calculate the minimum wavelength of light that would ionize a single electron from a single hydrogen atom. (b) In what part of the electromagnetic spectrum is a photon of this wavelength found? [4]

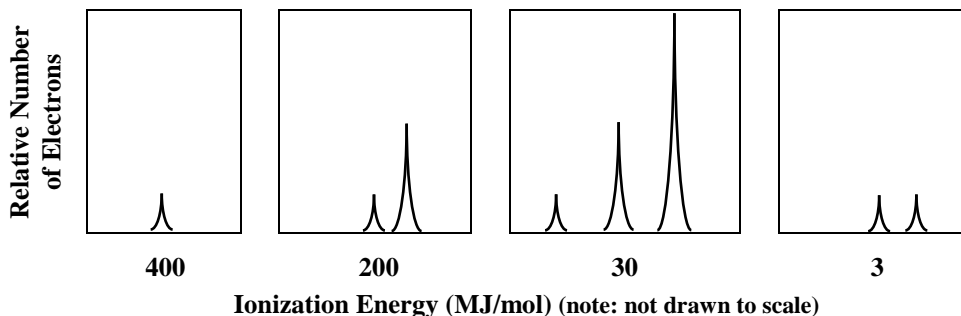
- 5) The ionization energy for the single electron in hydrogen is 1312 kJ/mol. Calculate the ionization energy for the single electron in hydrogen in eV. [3]

- 6) Identify the elements that compose the mixture, based on the bright-line emission spectra below. [1]

	400nm	500nm	600nm	700nm
Hydrogen				
Helium				
Lithium				
Sodium				
Mixture				

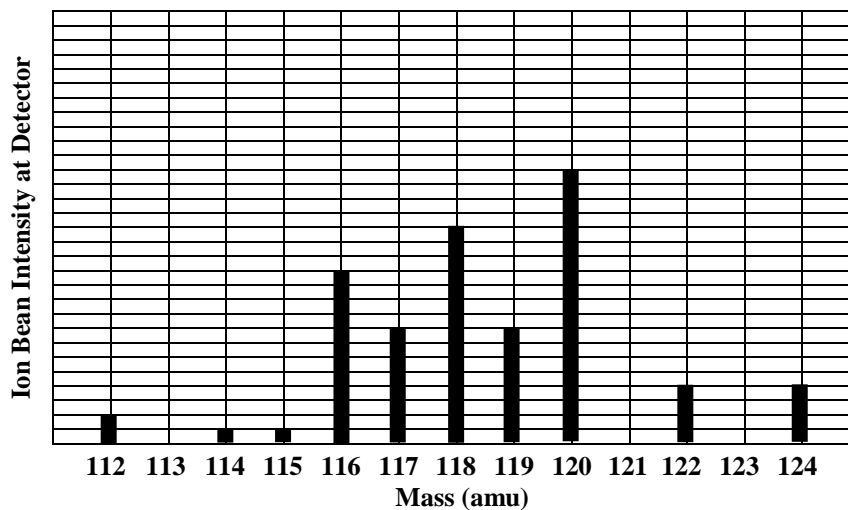
- (A) Sodium and Lithium
(B) Sodium and Helium
(C) Hydrogen and Helium
(D) Hydrogen and Lithium

- 7) Identify the element, based on its photoelectron spectrum below. [1]



- (A) Oxygen
(C) Titanium
(B) Calcium
(D) Germanium

- 8) Estimate the average atomic mass of the element, based on its mass spectrum (MS) below. [1]



- (A) 117 amu
(B) between 117 and 118 amu
(C) between 118 and 119 amu
(D) between 119 and 120 amu

- 9) Which of the following types of spectra can be used to determine which functional groups are present in a molecule due to the differences between energy states of bond vibration? [1]

- (A) A microwave spectrum
(B) An infrared spectrum
(C) An ultraviolet/visible absorption spectrum
(D) A nuclear magnetic resonance spectrum